Equilibrium Studies in Formylation Chemistry

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Executive Summary

The open literature for arene carbonylation was reviewed so as to permit a judgment on the suitability of further research and development into this potentially fruitful technology that might be applied for commercial benefit to the petroleum industry in the area of transportation fuels. Recent reports to BP(USA) are appended and discussed to shed light on the problem inherent to all acid-catalyzed, carbonylation reactions: the breaking of the acid/aldehyde complex.

Catalytic processes for arene carbonylation initiated by photons do exist however the yields are limited by the unfavorable thermodynamics for the reaction (~+5 kcal/mol). Radical-initiated methods have been reported which are not limited by equilibrium; however these routes require expensive reagents, such as potassium persulfate, to initiate the reaction and thus are not commercially viable.

The oldest process reported in the literature, the Gatterman-Koch process, requires aluminum chloride complex and HCl to initiate the reaction by protonation from a super acid and an additional aluminum chloride to sequester the product aldehyde so as to circumvent the equilibrium limitation. This process is further handicapped by the evolution of stoichiometric amounts of aluminum salt and HCl at the end of the process so as to release the product from the reaction mixture. As such, the environmentally aggressive process cannot be used for the intended reaction. The combination of HF/BF_3 is an alternative to the Gatterman-Koch process to produce arene aldehydes by an economically-feasible process; however, the use of HF represents an environmental threat that is unacceptable to most private companies doing business in the USA.

Trifluoromethanesulfonic acid is a possible replacement for HF/BF_3, which is a very efficient carbonylation agent that produces *p*-tolualdehyde from toluene in high yields (~95%). A closed-loop process has been reported for the production of *p*-tolualdehyde; however, the regeneration of the acid *required by this process* using trialkylamines is too expensive to be commercially-attractive for large-volume processes making commodity chemicals. The problem with this process is the large amounts of water needed to free the product aldehyde from the product mixture. In order for displacement to occur, the Gibbs energy of dissociated species must be lower than that of the complex. Appropriate additives yielding liquid/liquid mixtures should be such that one phase preferentially attracts the aldehyde while the other phase preferentially attracts the catalyst. Hydrogen bonding by additives may be used, but the aldehyde-rich phase

must exhibit low catalyst solubility. Any species that is more acidic than water may be appropriate; a highly polar liquid phase, immiscible with hydrocarbons, may be attractive for the catalyst. Small water additions may be appropriate if they are saturated with acidic salts.

One solution to this process may be the use of a co-solvent that can be recycled and which breaks the acid/aldehyde complex without the addition of water. We modeled this complex breaking phenomena using quantum mechanical methods to suggest that the acid/aldehyde complex may be broken by the use of ethers such as dimethyl or diphenyl ether, see Appendix II, June 2005 Progress Report to BP(USA). *If the results of these modeling efforts can be confirmed by laboratory testing*, then the trifluoromethanesulfonic acid process might be economically promising for the production of arene aldehydes from CO and arene.

Another approach to this problem is to use perfluoroalkanesulfonic acids having lower intrinsic acidities that allow the aldehyde to be vaporized from the product mixture with the addition of a co-solvent such as trifluoroacetic acid and without the addition of water to break the acid/aldehyde complex. This patented process claims that 1) the aldehyde can be recovered from the product mixture by vaporization, and that the trifluoroacetic acid can be distilled from the perfluoroalkanesulfonic acid. The patent inventors show that the reactivity of the perfluoroalkanesulfonic acids $[CF_3(CF_2)_nSO_3H]$ decrease with increasing values of n. This process must be reproduced by researchers in an independent laboratory so as to develop data necessary to complete an economic feasibility case for this process. This process may merit further attention if the perfluoroalkanesulfonic acids demonstrate acceptable reactivity, and if the product can be released from the product mixture by vaporization. Our attempts to model the breaking of the perfluoroalkanesulfonic acid/aldehyde complex suggest that other co-solvents such as dimethyl and diphenyl ether could be effective agents in breaking this acid/aldehyde complex, Appendix II, June 2005 report to BP(USA).

Solid acids derived from zeolites (X, Y, and ZSM-5) and ion-exchanged with Cu ions were reported to be effective arene carbonylation conversion agents with the addition of HCl gas along with CO. The surface of the acids becomes saturated with the product aldehydes and this surface must be regenerated before the solid can be reused by the addition of either steam, carbon dioxide or ammonia. In attempts to model the formation of surface aldehyde/surface acid sites, we modeled the zeolite surface as combined Brønsted/Lewis superacid. The structure of this site was developed from that of an ionic liquid containing the superacidic $Al_2Cl_7^-$ anion. This modeling showed that the superacid/aldehyde complex could be broken using either trimethylamine or tetramethylaminechloride as the complex breaking agents (Appendix II, March 2005 report to BP(USA)). Therefore, the use of an amine derivative is consistent with the patent report on the use of ammonia to break the complex.

This concept of carbonylation by fixed beds of zeolites requires that the zeolites be used in a cyclical mode much like adsorption beds to affect continuous conversion of the arene substrate. Other solid acids derived from sulfated zirconia have been reported

in the patent literature as effective arene carbonylation agents. Only small yields of the aldehyde (~1-3%) were observed when the surface became saturated with the product. No regeneration strategies were reported in this patent.

Acidic polymeric resins were reported as effective carbonylation agents in both the patent and refereed journal literature. We have attempted without success to reproduce these results in our laboratories. Even if we do become successful in using these acidic resins as carbonylation agents, there remains the problem of regeneration of the resin to recover the product aldehyde and to recover the reactivity of the solid. The thermal instability of polymeric resins precludes using higher temperatures to regenerate the solid.

Quo Vadis

As a result of this review of the literature it appears that three paths forward are possible: 1) investigate the efficacy of ethers in combination with other agents, such as trifluoroacetic acid or dichloromethane as a means to break the trifluoromethanesulfonic acid/aldheyde complex, 2) examine the perfluoroalkanesulfonic acid family as conversion agents for arene carbonylation (benzene or toluene) with an emphasis on the reuse of the sulfonic acid , 3) examine the reactivity and reuse of zeolites decorated with Cu and treated with HCl, and 4) high risk, high return technologies not yet proven nor patented. We will discuss each of these areas briefly here and attach a proposal (Appendix III) to elaborate further our ideas.

a. Complex breaking. Strong acids appear to be a viable path to forming the arene aldehydes in high yields by carbonylation of the substrate. However, a common problem for all systems is the breaking of the acid/aldehyde complex so as to regenerate the acid. Regeneration of the HF/BF_3 system is accomplished by partial vaporization of the HF from the reaction mixture followed by the introduction of the co-solvent and distillation of the resulting liquid to recover the product and regenerate the acid apart from the co-solvent. It is clear from our earlier work with triflic acid that heat alone will not break the acid/aldehyde complex. The results of modeling during the past 12 months suggest that agents such as dimethyl ether coupled with modest heating *in vacuo* might break the triflic acid/aldehyde complex. Other modeling results of superacidic ionic liquids show that both amines and ethers are viable complex-breaking agents. These modeling results appear to be confirmed by a recent manuscript submitted to MGW for review which showed that the combination of diethyl ether and dichloromethane were used to partially recover benzaldehyde from the product mixture. Finally, the theory of liquid-liquid extraction from ionic and non-ionic liquids suggest that additives to both the organic and inorganic phases are necessary to recover an organic agent bound up in the inorganic phase. We propose to examine the efficacy of a small number of extracting agents for recovering tolualdehyde from synthetic mixtures of triflic acid/tolualdehyde and

perfluoro-*n*-butanesulfonic acid/tolualdehyde. The list of extracting agents includes: dimethyl ether (under pressure), diethyl ether, diphenyl ether (elevated temperature, ~80°C), trimethyl amine, and tetramethylamine chloride.

b. Additional tests will be completed for breaking the perfluoro-*n*-butanesulfonic acid/tolualdehyde complex using trifluoroacetic acid as a co-solvent neat and in combinations with stoichiometric amounts of diethyl ether, or dichloromethane, or a mixture of dichloromethane/diethyl ether.

c. Further tests are proposed using biphasic liquid distillation as means to reduce the evaporation temperatures. Most commonly, steam injection is used to recover less volatile materials from a reaction mixture. In this case, we propose the use of dry perfluoro alkanes, which are typically immiscible with hydrocarbon mixtures.

2) Perfluoroalkanesulfonic acids. The patent literature (references 68-70 in citation list) suggests that one may use an acid of lower acidity than triflic acid to affect the arene carbonylation and also to recover the aldehyde with acid regeneration without the use of water to break the complex. We propose to reproduce these patent results using the protocol outlined in the patent examples and compare these results for aldehyde yields versus reaction time against the results obtained for triflic acid at the same reaction conditions. We anticipate that the yields obtained over pentafluoroethane sulfonic acid and perfluoro-*n*-butanesulfonic acid with be much lower than those obtained over triflic acid. The combination of this reaction activity work coupled with complex breaking activities in item (1) will determine if the perfluoroalkanesulfonic acids merit further examination in larger scale studies.

3) Solid acids. One patent report shows that arenes can be carbonylated over Cu-exchanged zeolites using a feed gas of vaporized arene, HCl gas, and CO at high pressure (~70 bar). Quite a bit of effort will be needed to confirm this patent if it is desired to complete the work in a three-phase, fixed bed, flow reactor. We have such a reactor available now, but we do not have a suitably-trained person to complete the work here. I anticipate that a technician or a post-doctoral fellow may be available at Mississippi State University when I move there in January. On the other hand, a simple reaction in a stirred autoclave will determine if the system demonstrates any activity and will tell us the regio-selectivity. We have the reactor and the ability to complete reactions in such an autoclave now. Regeneration and reuse of the catalyst will be difficult to do *in situ* without the purchase of an appropriate "catalyst basket" made by Parr for this autoclave, though such regeneration can be completed *ex situ* without the need for this device.

4) High risk, high return, untested technologies. The radical-initiated path to carbonylation is attractive in that no aldehyde complex is developed at the end of the process. However, the use of the radical initiator in stoichiometric proportions is commercially unattractive. We propose an alternative to this process which uses a highly reactive Ce metal oxide in acidic media [$H_2Ce(NO_3)_6$] to affect activation of the arene. The Ce metal ion cycles between

3+ and 4+ states to affect oxidation of the arene and the metal ion can be regenerated by electrochemical means. Moreover, ceric acid may be able to protonate both CO and the arene to form the formyl cation and the arene carbocation simultaneously so as to enhance the reaction rates. We propose to examine the viability of this concept to carbonylate toluene either to the aldehyde of possibly to the organic acid.

Appendix I—Detailed Literature Survey

Considerations in Examining the Literature

Some considerations were given in examining the literature. We did not include references that were written in a language other than English for which we did not have translations into English. Reactions which involved functionalized substrates, such as arene halides, were not included as the commercial value of this process will be limited since the substrate cost will be high. Also eliminated from further consideration were reactions involving formylating agents more expensive than CO. For example, Smith[1] describes a process where a mixture of hexamethylenetetramine and trifluoroacetic acid were reacted with arene in stoichiometric proportions to make an imine intermediate which was then converted to the aldehyde by the addition of water. The use of hexamethylenetetramine in stoichiometric proportions makes this process commercially unattractive.

Early Attempts at Arene Carbonylation using $AlCl_3$ Complex

The formylation of arenes has attracted the commercial and scientific interests of scientists and engineers for over a hundred years beginning with the publication of a manuscript by Gattermann and Koch in 1897.[2] In this early work aluminum chloride complex, copper chloride, HCl gas and CO were combined to produce benzaldehyde. Commercial interest for arene formylation continues to be high even though the one demonstrated process[3] uses the environmentally aggressive catalyst HF/BF_3. This process allows for the recovery of the aldehyde by breaking the aldehyde acid complex.[4] As a result of the desire to reduce the environmental threats, manufacturers seek agents for this conversion that poses minimal threats to the environment.

Photocatalytic Carbonylation by Metals

Other approaches for this reaction include the oxidative addition to the metal center in a suitable complex usually accompanied by irradiation. Such metal complexes include: $IrH_3(CO)(Ph_2PCH_2CH_2PPh_2)$;[5] $RhCl(CO)(PPh_3)_2$;[6] and $RhCl(CO)(PMe_2)_2$.[7] Gordon and Eisenberg report the UV-photocatalysis of benzene to benzaldehyde using Ru(0) complexes[8], whereas Kunin and Eisenberg used either a Rh(I) or an Ir(I) complex.[9] These authors confirm the low yields of the aldehyde (~5%) when equilibrium thermodynamics (ΔG°_{rxn} = 1.7 kcal/mol arene reacted)[9] of the neutral, uncomplexed species control the system. In a separate experiment, they observed decomposition of the aldehyde product to CO and benzene, confirming that the reaction is catalytic and that it is equilibrium controlled. They speculate that UV photons are used to activate a C-H bond and that the metal complex is involved in adding the CO to the activated phenyl-H species. This photocatalytic approach will never realize high yields as long as the product and reactants coexist with a viable catalyst in the presence of photons. They attempted, without success, to drive the reaction to completion by 1) hydrogenating the aldehyde to the alcohol, and by 2) trapping the product by hydrosilation *in situ* with the photocatalysts. The authors could not explain their lack of success in trapping the aldehyde as another product of lower free energy but it is clear this approach does work when the product aldehyde forms an adduct with a strong acid followed by destruction of the catalyst by the action of water (*vide infra*). Even if they were successful in driving the reaction; the turnover rate, 0.4 h^{-1}, is too low to be of commercial importance even though the conditions were modest (~25°C, 800 Torr CO). It appears that the thermodynamic limitation does not permit sufficient thermodynamic driving force to develop suitable reaction velocities.

Radical-Initiated Catalytic Carbonylation by Metals

Zhuo and Jiang reported a different system for the catalytic carbonylation of benzene using $Co(OAc)_2/K_2S_2O_8/CCl_3COOH$/pyridine which realized yields of benzaldehyde up to 35% when the benzene conversion was 50%.[10] It was speculated that a radical mechanism was operative as a result of either a radical formed from either the persulfate or the trichloroacetic acid. The use of either required reagent makes this scheme commercially unfeasible. Moreover, the oxidation of the methyl carbon of toluene may lead to another reaction manifold as reported by George, *et al.*[11] Soma[12] reported in a Japanese Patent Application the production of aromatic aldehyde by the action of Cu(I) or Ag(I) carbonyls

in super acids having a Hammett acidity function of –15 and more negative. In view of the reports to follow, it appears that the incorporation of the metal salts are immaterial to the aldehyde production and that the super acids alone are responsible for the high yields of aldehyde.

Perfluoroalkanesulfonic Acids

[a]"Perfluoroalkanesulfonic acids and their derivatives are of commercial significance because of their unusual acid strength, chemical stability, and the surface activity of the higher members of the series (eight carbons and larger). The best method of preparation is *via* the perfluorinated sulfonyl fluorides derived from electrochemical fluorination, although other methods including direct fluorination with fluorine gas are known. The first member of the series, trifluoromethanesulfonic acid, is one of the strongest acids known. Alkyl esters of trifluoromethanesulfonic acid are among the best leaving groups known and are commonly used in that capacity in displacement reactions. Lithium trifluoromethanesulfonate solutions are very conductive and can find application as electrolytes in primary lithium batteries and as antistats for films."

"Derivatives of the longer-chain perfluoroalkanesulfonic acids have a number of unique surface-active properties and have formed a basis for a number of commercial products. The longer-chain acids and their salts, particularly $C_8F_{17}SO_3H$ and higher, are surface-active agents in aqueous media. They reduce the surface tension of water to levels not possible with hydrocarbon surfactants."

"1. Preparation. The perfluoroalkane sulfonic acids were first reported in 1954. Trifluoromethanesulfonic acid was obtained by the oxidation of bis(trifluoromethyl thio) mercury with aqueous hydrogen peroxide[13]. The preparation of a series of perfluoroalkanesulfonic acids derived from electrochemical fluorination (ECF) of alkane sulfonyl halides was also disclosed in the same year[14]. The synthetic operations employed when the perfluoroalkanesulfonic acid is derived from electrochemical fluorination, which is the best method of preparation, are shown in the equations that follow[13,14,15]"

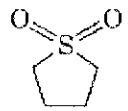
$$R_hSO_2F + HF \rightarrow R_fSO_2F + H_2 \quad (1)$$

where R_h is an alkyl group and R_f is a perfluoroalkyl group

$$R_fSO_2F + KOH \rightarrow R_fSO_3K + HF \quad (2)$$

$$R_fSO_3K + H_2SO_4 \rightarrow R_fSO_3H + KHSO_4 \quad (3)$$

"Perfluorosulfonyl fluorides can also be prepared by the electrochemical fluorination of saturated or unsaturated cyclic sulfones [15,16,17]. Perfluorobutanesulfonyl fluoride can be prepared in 40–48% yield from sulfolane (eq. 4)[18].



$$\xrightarrow{ECF} CF_3CF_2CF_2CF_2SO_2F \quad (4)$$

"Yields of sulfonyl fluorides prepared by ECF vary depending on the particular structure. Chain degradation becomes more important as the chain length increases[18]. Yields can vary from 96% for perfluoromethanesulfonyl fluoride[19] to 43–50% for perfluorooctanesulfonyl fluoride[20]."

"Other preparations of trifluoromethanesulfonic acid include oxidation of methyltrifluoromethyl sulfide under a variety of conditions[21,22]. Perfluorosulfonyl fluorides have also been prepared by reaction of

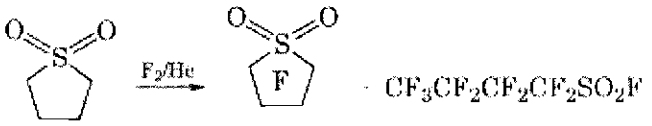
[a] Text in quotes were taken *verbatim* from the report by Patricia Savu, "Fluorine-Containing Polymers, Perfluoroalkanesulfonic Acids", *Kirk-Othmer Encyclopedia of Chemical Technology* (1994), John Wiley & Sons, Inc.

fluoroolefins with sulfuryl fluoride[23,24]. Chinese chemists have published numerous papers on the conversion of telomer-based alkyl iodides to sulfonyl fluorides[25,26] (eqs. 5 and 6):

$$2\ R_fI + Na_2S_2O_4 + NaHCO_3 \rightarrow 2\ R_fSO_2Na \quad (5)$$

$$R_fSO_2Na + Cl_2 \rightarrow R_fSO_2Cl + NaCl \quad (6)$$

"Perfluorosulfonyl fluorides have also been prepared by direct fluorination, although in general yields are lower than preparation by ECF. Perfluoromethanesulfonyl fluoride has been produced in 15% yield from direct fluorination of dimethyl sulfone[27]. Perfluoro-2-propanesulfonyl fluoride was prepared in 29% yield from propanesulfonyl fluoride[28]. Direct fluorination of tetramethylene sulfone leads to the intact perfluorinated sulfone in 28% yield and the ring-opened product (perfluorobutanesulfonyl fluoride) in 10% yield (eq. 10)[23].



(7)

"Currently, the commercially important methods of preparations of perfluorinated sulfonic acid derivatives are electrochemical fluorination and sulfur trioxide addition to tetrafluoroethylene with subsequent ring opening. The boiling points of a series of perfluoroalkanesulfonic acids are listed in Table 1[13].

Table 1. Boiling Points of Perfluoroalkanesulfonic Acids

Compound	CAS Registry Number	Bp, °C/kPa[a]	Bp, °C[b]
CF_3SO_3H	[1493-13-6]	60/0.4	166
$C_2F_5SO_3H$	[354-88-1]	81/2.9	175[c]
$C_4F_9SO_3H$	[59933-66-3]	76–84/0.017	200[c]
$C_5F_{11}SO_3H$	[3872-25-1]	110/0.67[d]	212[c, e]
$C_6F_{13}SO_3H$	[355-46-4]	95/0.47	225[c]
$C_8F_{17}SO_3H$	[1763-23-1]	133/0.8	249
4-CF_3(cyclo-C_6F_{10})SO_3H	[374-62-9][f]	120/0.4	241
4-C_2F_5(cyclo-C_6F_{10})SO_3H	[335-24-0][f]		254

[a]To convert kPa to mm Hg, multiply by 7.5; [b]At ; [c]Estimated; [d]The hydrate; [e]$C_5F_{11}SO_3H$ anhydrous. [f]Potassium salt.

Trifluoromethanesulfonic Acid

"The first member of the series, CF_3SO_3H , has been extensively studied. Trifluoromethanesulfonic acid [1493-13-6] is a stable, hydroscopic liquid, which fumes in air. Addition of an equimolar amount of water to the acid results in a stable, distillable monohydrate, mp 34°C, bp 96°C at 0.13 kPa (1 mm Hg)[29]. Measurement of conductivity of strong acids in acetic acid has shown the acid to be one of the strongest protic acids known, similar to fluorosulfonic and perchloric acid[30]."

"Trifluoromethanesulfonic acid is miscible in all proportions with water and is soluble in many polar organic solvents such as dimethylformamide, dimethylsulfoxide, and acetonitrile. In addition, it is soluble in alcohols, ketones, ethers, and esters, but these generally are not suitably inert solvents. The acid reacts

with ethyl ether to give a colorless, liquid oxonium complex, which on further heating gives the ethyl ester and ethylene. Reaction with ethanol gives the ester, but in addition dehydration and ether formation occurs."

"Alkyl esters of trifluoromethanesulfonic acid, commonly called triflates, have been prepared from the silver salt and an alkyl iodide, or by reaction of the anhydride with an alcohol[30,31,32]. Triflates of the 1,1-dihydroperfluoroalkanols, $CF_3SO_2OCH_2R_f$, can be prepared by the reaction of perfluoromethanesulfonyl fluoride with the dihydroalcohol in the presence of triethylamine[33,34]. Triflates are important intermediates in synthetic chemistry. They are among the best leaving groups known, so they are commonly employed in anionic displacement reactions."

"The metallic salts of trifluoromethanesulfonic acid can be prepared by reaction of the acid with the corresponding hydroxide or carbonate or by reaction of sulfonyl fluoride with the corresponding hydroxide. The salts are hydroscopic but can be dehydrated at 100°C under vacuum. The sodium salt has a melting point of 248°C and decomposes at 425°C. The lithium salt of trifluoromethanesulfonic acid [33454-82-9], CF_3SO_3Li , commonly called lithium triflate, is used as a battery electrolyte in primary lithium batteries because solutions of it exhibit high electrical conductivity, and because of the compound's low toxicity and excellent chemical stability. It melts at 423°C and decomposes at 430°C. It is quite soluble in polar organic solvents and water. Table 2 shows the electrical conductivities of lithium triflate in comparison with other lithium electrolytes which are much more toxic[35]."

Table 2. Comparative Electrical Conductivity[a] of Lithium Salts

Concentration, *M*	CF_3SO_3Li	$LiClO_4$	$LiAsF_6$
0.5	24.4	29.9	26.3
0.1	5.81	7.4	6.94
0.05	3.12	3.83	3.57
0.01	0.70	0.83	0.79

"In water at 25°C,

"Due to the strong ionic nature of lithium trifluoromethanesulfonate, it can increase the conductivity of coating formulations, and thereby enhance the dissipation of static electricity in nonconducting substrates (see Antistatic agents)[36]."

"Trifluoromethanesulfonic acid anhydride, bp 84°C, is prepared by refluxing the acid over an excess of phosphorous pentoxide[30,37]. The anhydride reacts instantaneously with ammonia or amines to form trifluoromethanesulfonamides. The anhydride reacts with most polar organic solvents. It polymermizes tetrahydrofuran, THF, to give a living polyether having cationic activity at each chain end[38]."

"Several review articles[39,40,41,42,43] cover the chemistry of the acid and its derivatives in great detail. Trifluoromethanesulfonic acid is available from the 3M Co. as Fluorochemical Acid FC-24; the lithium salt is available as Fluorochemical Specialties FC-122, FC-123, and FC-124[44]."

Solubility of *p*-tolualdehyde in triflic acid/water mixtures. We examined the effect of water to break the triflic acid-aldehyde complex by preparing an acid/aldehyde sample having a composition similar to the product at the end of the carbonylation reaction: 10 moles acid/mole of aldehyde by adding 1.95 g of triflic acid with 0.16 g of aldehyde. To this mixture we added distilled water and recorded the color of the solution and the number of phases present. The liquid remains red for a water/HOTf ratio up to 7 mols/mol. When the amount of water exceeds 8 mols/mol acid, the solution turns a very dark green, but it remains one phase. Phase separation was observed when the amount of water equals 13 mol water/mol of acid. Further addition of water makes the thin, green oil layer more obvious, but the lower layer is a light

green color. After 1 h the phase separation is very distinct with the lower phase showing a very light green color and the upper phase showing a very dark, green color. This liquid is still very acidic as shown by Litmus paper. The lower phase becomes colorless after standing for 24 h.

Solubility of toluene in triflic acid. We determined the solubility of toluene in triflic acid at room temperature by the following method. 0.5 mL of triflic acid was dispensed into a 2 mL glass sample bottle. To this acid was added toluene in 100 μL aliquots. We observed the system as a function of toluene addition. The first 100 μL of toluene (0.94 μmoles) was completely soluble in the 0.5 mL of triflic acid (5 mmoles); however, the next 100 μL of toluene was only partially dissolved as we noticed the formation of a second layer in the sample vial. Three more 100 μL aliquots of toluene when added to the acid-rich phase did not dissolve. Thus, the solubility of toluene in triflic acid at room temperature is less than 1.8 mmol of toluene per 5 mmoles of triflic acid for a ratio of 0.376 mol toluene/mol of triflic acid (toluene mole fraction = 0.27). To this mixture we added 100 μL of *p*-tolualdehyde (0.91 mmole). The resulting red liquid was a single phase. Apparently, the addition of the *p*-tolualdehyde to the acid/toluene mixture increases the solubility of the toluene in the acid to a ratio of 0.94 mol toluene/mol acid. Subsequently, we added two more aliquots of toluene, each 0.94 mmoles, to this mixture of acid/toluene/tolualdehyde. The resulting solution again showed two phases when the toluene mole fraction was > 0.44. This phase tested acidic by litmus paper. We continued this study by adding another aliquot of *p*-tolualdehyde (0.91 mmol, total = 1.82 mmol) to this solution to observe that again, only a single phase was present. We continued to add toluene in 100 μL aliquots until a second phase was formed when total amount of toluene was > 9.42 mmol. Then a third aliquot of *p*-tolualdehyde was added (2.73 mmol, total) and more toluene was added until phase separation occurred (14.79 mmol of toluene). Finally, we added a fourth aliquot of p-tolualdehyde (total amount of *p*-tolualdehyde = 3.64 mmol) to this mixture and again a single phase was observed. To this solution was added more aliquots of toluene (total amount added = 7.86 g toluene) and no phase separation was observed. We conclude that toluene is miscible with a solution containing 3.6 mmol of *p*-tolualdehyde with 5.0 mmol of triflic acid.

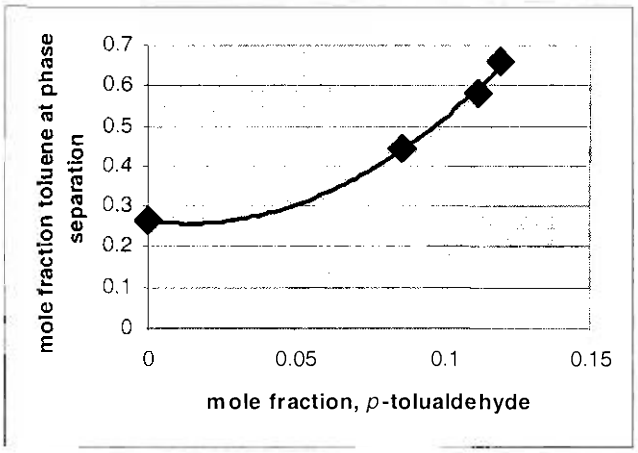


Figure 1 Solubility of toluene in triflic acid

The solubility of toluene in the acid/*p*-tolualdehyde/toluene solution is a function of *p*-tolualdehyde mole fraction in Figure 1. Triflic acid dissolved toluene up to a toluene mole fraction = 0.27 at phase separation. With the addition of 0.91 mmol *p*-tolualdehyde, the toluene solubility increased to a mole fraction of 0.44 at phase separation. Upon adding another 0.91 mmol *p*-tolualdehyde to this solution, the toluene mole fraction at phase separation increased to 0.55 and thence to 0.66 with the last aliquot of *p*-tolualdehyde (*p*-tolualdehyde mole fraction = 0.12). These data suggest that the solubility of toluene in triflic acid can be improved by the addition of *p*-tolualdehyde.

Enthalpy of mixing-binary mixtures with HOTf and base. The enthalpy of mixing, Figure 2, was estimated for combining triflic acid with the following bases: 1) water, 2) *p*-tolualdehyde, 3) methyl pyrrolidine, 4) acetonitrile, and 5) tetrahydrofuran. The enthalpy of mixing sulfuric acid with water was also estimated using this apparatus as a control. Our measured values for the enthalpy of mixing for sulfuric acid/water are similar to those found in the literature and thereby show the accuracy of our apparatus (+/- 7 kJ/mol = 1-2 kcal/mol). We notice that the enthalpy of mixing triflic acid/water (40-50 kJ/mol of acid) is greater than that for sulfuric acid/water by about 10-15 kJ/mol acid (25-40 kJ/mol acid) when measured at the same ratio of base added to acid. However, the enthalpy of mixing of *p*-tolualdehyde is less than the heat of mixing of triflic acid/water (25-30 kJ/mol). Thus, we see why water can be used to "break" the aldehyde/acid adduct at the end of reaction.

The heats of mixing are even higher for combining triflic acid with other bases, such as methyl pyrollidine (53-60 kJ/mol acid), tetrahydrofuran (43 kJ/mol acid), and acetonitrile (42 kJ/mol acid). Methyl pyrollidine (MPy) is used in a patented process[45] to dry wet triflic acid; thus from these thermodynamic data, it becomes easy to see why it can displace water from the wet acid. This base will form two salts with triflic acid which engage in the following equilibrium upon addition/subtraction of heat:

$$\text{MPy(HOTf)}_2 \overset{+\Delta}{\underset{-\Delta}{\Leftrightarrow}} \text{MPy(HOTf)} + \text{HOTf(g)}$$

Thus, HOTf is freed from the diacid salt of MPy by the addition of heat so as to vaporize the acid and to leave the monoacid as a liquid in the still. The monoacid, MPy(HOTf), is sufficiently basic to displace water from wet HOTf by the following equilibrium.

$$\text{MPy(HOTf)} + \text{HOTf}(n\text{H}_2\text{O}) \overset{+\Delta}{\underset{-\Delta}{\Leftrightarrow}} \text{MPy(HOTf)}_2 + n\text{H}_2\text{O(g)}$$

Water is eliminated from the wet acid at 30°C under a vacuum of 1 Torr with the action of the monoacid. Thus, the Haldor-Topsoe process dries wet triflic acid by the combination of these two equilibria which are driven by the combination of heat and vacuum.

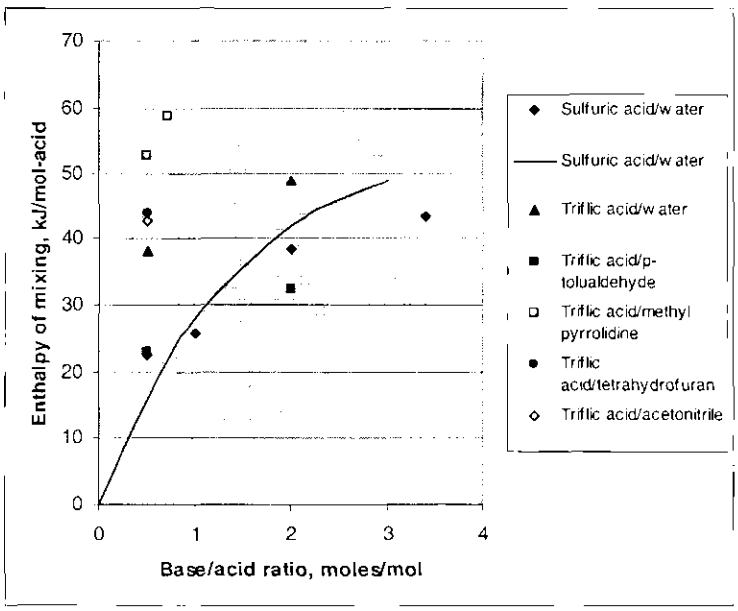


Figure 2 Enthalpy of mixing for acids/bases

We tested tetrahydrafuran and acetonitrile as alternatives to MPy as a co-solvent to break the aldehyde acid adduct.

Heats of mixing—ternary mixtures. We examined further the heat effects for combining HOTf with water and organic bases such as methyl pyrrolidine and also p-tolualdehyde. In the first instance we estimated the heat effect for combining a mixture containing 2 g-mol of HOTf and 1 g-mol of p-tolualdehyde with 4 g-moles of water. The effect was -34.5 kJ/mol of acid. In another experiment, we combined a mixture containing 1.5 g-mol of HOTf and 1 g-mol of MPy with another mixture containing 1 g-mol of HOTf and 2 g-mol of water. The resulting heat effect was -5.8 kJ/mol acid. These heats of mixing are smaller than what we reported for mixing of HOTf with a single base by a factor of 2-10; perhaps as a result that both of the reagents were "diluted" HOTf mixtures before they were combined.

Reactivity of Triflic Acid Towards Arene Carbonylation

Effect of stirring speed. Tests were completed to identify the stirring speed necessary to eliminate the effects of mass transport upon the observed reaction rates. The impeller speeds were varied between 500 and 2000 rpm for a constant set of reaction conditions. For all tests, the reaction temperature was 25°C, the reaction time was 30 minutes, and the acid/substrate ratio was 2 mol/mol. In one series of tests the CO partial pressure was 809 psig and for other series the CO partial pressure was 1100 psig. The conversion of toluene was recorded for each of these tests as an indication of the reactor performance (Table 1).

Table 1 Toluene conversion as a function of impeller speed

Impeller speeds, rpm	500	600	1000	1300	2000
CO partial pressure, psig					
809	3.8	6.5	8.0	--	8.0
1100	--	20.	--	20.3	20.6

These data suggest that the reaction rate was not affected by impeller speed when it was > 1000 rpm at > 800 psig pressure or when it was ≥ 600 at 1100 psig. For the remainder of the tests, the impeller speed was 600 rpm.

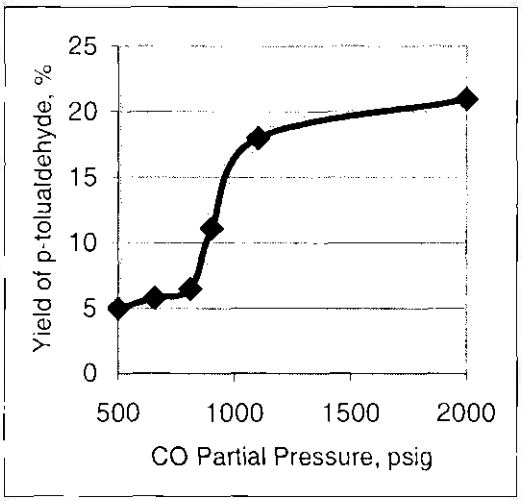


Figure 3 Effect of pressure upon reactivity

Effect of CO pressure. The CO partial pressure was changed from 500 to 2000 psig to document the effect of CO partial pressure on the reaction rate (Figure 3). For all tests, the acid/substrate ratio was 2 mol/mol, the reaction temperature was 25°C, and the reaction time was ½ h. The yields of *p*-tolualdehyde increased from 5-6% to 18% when the CO partial pressure increased from 800 to 1100 psig. Further increases in CO partial pressure to 2000 psig caused only small increases in the yield of *p*-tolualdehyde (21%). These tests suggest that the system becomes saturated with CO at a pressure near 1100 psig and therefore, the solubility of CO into the reacting liquid will not limit the reaction rate when the CO partial pressure is at least 1100 psig.

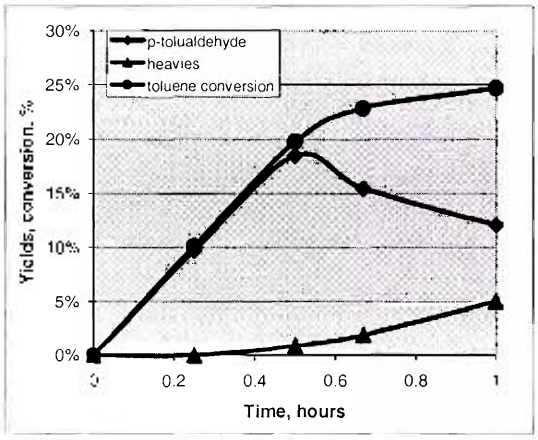


Figure 4 Reaction conversion *versus* time

Effect of reaction time. The kinetics of the reaction were measured at a CO partial pressure of 1100 psig and an acid/substrate ratio of 2 mol/mol at 25°C (Figure 4). For reaction times < ½ h, the selectivity to *p*-tolualdehyde was > 90% with *o*-tolualdehyde and heavier products (di- and tritolymethane) as the only other products that we observed. For longer reaction times, the heavier products appear in combined yields up to 5%. When the reaction time was 17.5 h, the yield of heavier products was 14.6%; whereas, the yield of o + p-tolualdehyde was 13.9%. The appearance of the heavier products was expected since Olah[46] and co-workers observed these products when toluene was mixed with *p*-tolualdehyde in the presence of triflic acid at room temperature.

We attempted to fit the toluene conversion data using irreversible, first-order kinetics (Fig. 5). These data could be fit using only first-order kinetics during the first ½ h of reaction, but the data appeared to depart from first order behavior when the reaction time was greater than ½ h. The departure from first-order kinetics coincided with the consumption of *p*-tolualdehyde by a secondary reaction to form di- and tritolymethane (Figure 3). This result suggests that the reaction sequence is more complicated than what can be modeled by simple, first-order kinetics. Olah showed that *p*-tolualdehyde can react with toluene in triflic acid to form di-and tri-tolymethane.[46] This reaction will consume *p*-tolualdehyde and we expect that the concentration of the aldehyde should pass through a maximum, as we observe. The rate to consume toluene by these two sequential reactions apparently slows as ditolymethane is formed and as the concentrations of both toluene and *p*-tolualdehyde fall. Thus, the apparent order changes as the reaction proceeds to form the di- and tri-tolymethanes.

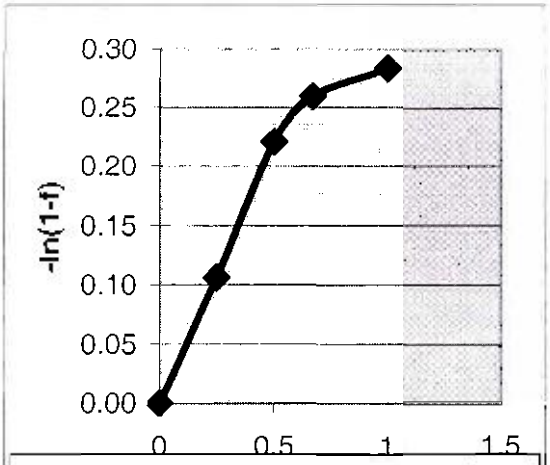


Figure 5 Fit of Toluene Conversion data to a First-Order Plot

Temperature effects. We studied the effect of reaction temperature in separate tests at 2, 25, and 50°C at a constant acid/substrate ratio of 2 mol/mol and for a reaction time of ½ h at a CO partial pressure of 1100 psig. These results are shown in Table **2**. Increasing the temperature from 2 to 50°C resulted in an increase in conversion with an optimum yield of p-tolualdehyde at 25°C (18.5%). Apparently, increased reaction temperatures favor the formation of the heavier products at the expense of tolualdehyde.

Reactivity of toluene carbonylation reaction in mixtures of triflic acid and base. We examined the room temperature reactivity of the toluene carbonylation reaction in a mixture of triflic acid and a base, with the base being either water or *p*-tolualdehyde (Figure 6). The initial starting mixture contained 2 moles of triflic acid, 1 mol of toluene and x moles of added base, either water or *p*-tolualdehyde. The total pressure was 150 psig after the CO was added and the room-temperature, reaction was stopped after ½ h. At these conditions, toluene is only partly soluble in triflic acid when no *p*-tolualdehyde was added.

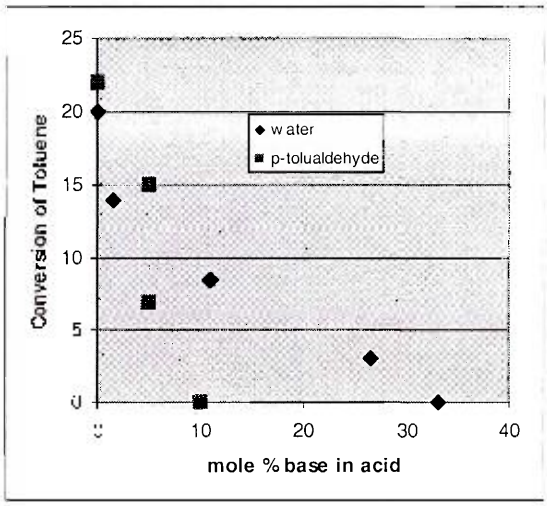


Figure 6 Toluene reactivity in triflic acid and added base

However, when the mole fraction of *p*-tolualdehyde was 0.1, all of the toluene was soluble in the initial mixture of triflic acid, but the reaction mixture was inactive when this much *p*-tolualdehyde had been added. When the initial mixture contained 5 mol% *p*-tolualdehyde, the conversion of toluene was between 8 and 15% in replicate tests.

Mixtures of water and triflic acid formed the hemi-hydrate and monohydrate which are both solids at room temperature and are totally unreactive towards the carbonylation reaction. Systematic decreases in the toluene conversion after ½ h was noted for increasing amounts of water initially from 2 mol% to 28 mol% (Fig. 6). The effect of adding the base water to the reactivity of the system was different from that of adding *p*-tolualdehyde, perhaps as a result of increased toluene solubility for small incremental amounts of *p*-tolualdehyde. The "lethal" dose of *p*-tolualdehyde was 10 mol% whereas the lethal dose of adding water was 33%. This result appears to be contra-intuitive from a consideration of the enthalpy of mixing data where the heat of mixing of water with the acid is much higher than the heat of mixing of the aldehyde with the acid. Molecular modeling of the respective acid/base adducts may shed light on this issue.

Table 2 Effect of Reaction Temperature

Temperature, °C	Conversion, %	Yield, % *p*-tolualdehyde
2	12.0	9.5
25	19.8	18.5
50	25.2	15.2

Effect of acid/substrate ratio.. It has been demonstrated[47] that increasing the acid/substrate ratio produced a beneficial effect upon yields of other hydrocarbon conversions. We examined this effect in separate tests for which we increased the acid/substrate ratio from 1-14 mol/mol when all other reaction conditions were

held constant: temperature = 25°C, reaction time = ½ h, CO partial pressure = 1100 psig. The data of toluene conversion and yield of *p*-tolualdehyde appear to correlate with the molar ratio of acid/substrate (Figure 7). These data showed that the conversion and yield increased with increasing acid/substrate ratio and that high conversion of the toluene was possible (~89%) at rather short reaction times, ½ h, when the acid/substrate ratio was >10. Moreover, the reaction was selective to the *para* isomer (~94% *p*-, 0.4% *m*-, and 5.5% *o*-tolualdehyde) at high acid/substrate molar ratios (10-14).

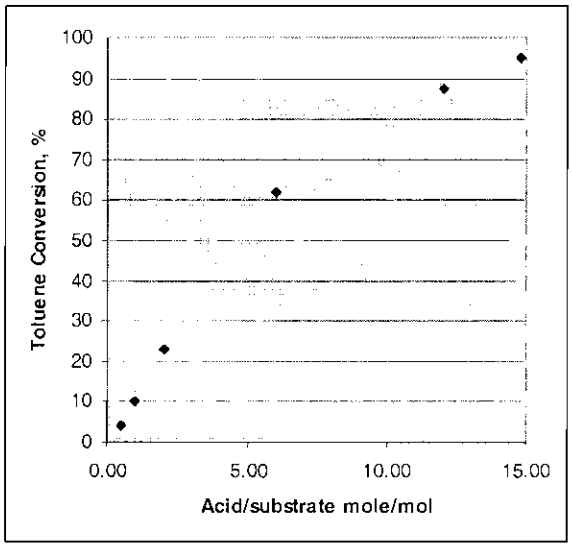


Figure 7 Effect of acid/substrate ratio on conversion

Data for acid/substrate ratios < 6 form a line for which the conversion appears to be $1/10^{TH}$ the value of the acid/substrate ratio. We may interpret these results by assuming that one mole of acid is required to protonate one mole of CO molecules to make the formyl cations. These cations attach to one mole of toluene to form the tolualdehyde carbocation. The remaining acid molecules are used to stabilize the tolualdehyde carbocations and to protonate the unreacted toluene so that the unfavorable side reactions do not occur between $[H\text{-toluene}]^+$ and the tolualdehyde carbocation. We speculate that the excess acid prevents the formation of di- and tri-tolymethane. Data at the high conversion do not fall along the line as a result of mass transport limitations on CO solubility in the liquid and/or as a result of reaction kinetics limiting the reaction rate at these high toluene conversions.

The response of this system was similar to that observed by Lin and MacCaulay[47] for the isomerization of trimethylbenzenes at 100°C in excess HF (7 mol HF/mol substrate) and for which the amount of BF_3 was varied. When the BF_3/substrate ratio was less than unity, the mesitylene content increased linearly with increasing amount of BF_3, and when the BF_3/substrate ratio was greater than unity, the mesitylene yield was 100%. They attributed the high yields of mesitylene to the super acidity of a mixture of BF_3 with HF which stabilized the BF_3/HF/mesitylene complexes. Similar results were reported for the isomerization of xylenes and tetramethylbezenes in excess superacid.[48] In these studies, the *meta* isomer was favored at high acid/substrate ratios. Olah[48] reported similar results for the isomerization of di-*iso*propylbenzene in excess super acid (either HF or triflic acid). Our recent work[49] described how high acid/substrate ratios, 20-50 mol/mol, favored the isomerization of dimethylbiphenyl to the *meta* isomers (*i. e.*, 3,3'- and 3,4'-dimethylbiphenyl) in preference to the other four isomers. The present data showed high yields of *p*-tolualdehyde at high acid/substrate and we suggest that the high yield/selectivity to the *p*-tolualdehyde could be attributed to the formation of a stable HOTf/*p*-tolualdehyde complex or some intermediate that leads ultimately to *p*-tolualdehyde.

Effect of substitution on the ring. We carbonylated *p*-xylene, 1,3,5-trimethylbenzene, and 1,2,4-trimethylbenzene in triflic acid (10 mol acid/mol substrate) at room temperature for 30 minutes using pressurized CO at 1100 psig. The purpose of these tests was to determine the effect of ring substitution on the reaction rates and the selectivity of the carbonylation. The results are summarized in Table **3**.

When the substrate was toluene, two products were observed: *p*-tolualdehyde and *o*-tolualdehyde. The regioselectivity to the *para*-isomer was 94% and the remaining was the *ortho* isomer (5%) with only a small amount of the *m*-isomer (0.5%), and the conversion of toluene was 99%.

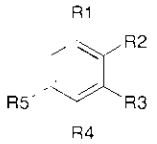
The effect of adding a second methyl group to the ring was investigated by using *p*-xylene as the substrate. At room temperature, where all of the ingredients were added together and reacted for ½ h, only two major products were observed: 2,5-dimethylbenzaldehyde, selectivity = 97% and 2,4-dimethylbenzaldehyde, selectivity = 3%. These results suggest that the isomerization reaction was slow compared to the carbonylation reaction at room temperature. A follow-up reaction was completed in which all of the same reagents were present initially but the vessel was not pressurized with CO. The result of this

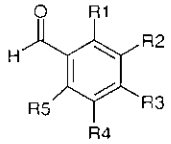
experiment showed that very little of the *para*-xylene was isomerized at room temperature in ½ h of reaction time. However, McCaulay and Lien[47] showed that all *p*-xylene was isomerized to *m*-xylene at the higher temperature of 80 °C and ½ h, so that elevated reaction temperatures are necessary to affect the isomerization of the xylenes in HOTf.

We examined further the effect of adding methyl groups to the ring by completing two runs using the following two substrates: 1,3,5-trimethylbenzene and 1,2,4-trimethylbenzene. Only a single one regioisomer was formed using mesitylene as the substrate: 2,4,6-trimethylbenzaldehyde with a small amount of another product having a m/e = 202. The conversion of the mesitylene was 71% in this test. When the reactant was 1,2,4-trimethylbenzene, the conversion was 99% and two products were found: 2,3,5-trimethylbenzaldehyde (94% selectivity) and 2,4,6-trimethylbenzaldehyde (6% selectivity). Apparently, a small portion of the substrate isomerized during the reaction. It is surprising that the reactivity of the mesitylene was lower than that observed for the 1,2,4-trimethylbenzene (71% conversion vs. 99%) and that the conversion of *p*-xylene was lower than that for toluene (82% vs. 99% for toluene). We have found that mesitylene, when added as a solvent, completely inhibited the acid-catalyzed isomerization of dimethylbiphenyl (dmbp) and that *p*-xylene, when added as a solvent, significantly slowed the isomerization rate.[49] Apparently, these two substituted arenes, *p*-xylene and mesitylene, inhibited the dmbp isomerization reaction; perhaps, by combining with HOTf to form stable carbocations and thus lower the overall acidity of the system.

Table 3 Selectivities/Conversion for Carbonylation of Substituted Benzenes at 298 K

Substrate Products





R1	R2	R3	R4	R5	Conversion	Selectivity	
H	H	Me	H	H	99	94 (R3 = Me);	5 (R1 = Me)
Me	H	H	Me	H	82	97 (R1=R4= Me);	3 (R1=R3 =Me)
Me	H	Me	H	Me	71	100 (R1=R3=R5=Me).	
Me	Me	H	Me	H	99	94 (R1=R2=R4=Me)	6 (R1=R3=R5=Me)

Predictions—Enthalpy of mixing of HOTf binary mixtures. We attempted to model the enthalpy of mixing using AM1 to determine the enthalpy of formation for the pure components and for the mixtures that had been subjected to energy minimization to obtain the optimum geometry. Thermodynamic principles were used to express the heat of mixing in terms of measured temperature rise in the calorimeter and the heat capacities of the components of the mixture. Consider the definition of the heat of mixing for a non-ideal mixture:

$$\Delta H^{\circ}_{mixture} = \Sigma \Delta H^{\circ}_{i} + \Delta H^{\circ}_{mixing}$$

where, $\Delta H^{\circ}_{mixture}$ = enthalpy of the entire mixture,
ΔH°_{i} = enthalpy of pure component i
$\Delta H^{\circ}_{mixing}$ = heat of mixing for this mixture

The quantum mechanics programs predict the properties of the species as an isolated molecule and thus will not model solvent interactions. We may include this consideration by modeling ensembles of the species so as to estimate the enthalpies of these interactions. Consider the data in table 4 where we show the enthalpy of formation for tolualdehyde, water or triflic acid molecules in ensembles numbering from 1 to 12 molecules. Notice how the enthalpy of formation for the isolated molecules is less negative than that predicted for the same molecule when surrounded by like neighbors. That is, the standard state, enthalpy of

formation for water as an ideal gas molecule is estimated to be -59.24 kcal/mol; whereas, the reported enthalpy of formation for gaseous water at 298 K is -57.796 kcal/mol.[50] For larger ensembles of water, the AM1 software predicts more negative, enthalpies of formation up to a value of -67.34 kcal/mol. The heat of formation of liquid water at this condition is -68.32 kcal/mol.[50] The increased negative, enthalpy of formation is due to solvent effects that appear to be modeled by AM1.

Table 4 Enthalpy of formation for ensembles of pure compounds, kcal/mol

n	1	2	3	4	5	8	10	12
water	-59.24	-61.74	-64.01	-65.22	-65.72	-66.47	-66.88	-67.34
p-tolualdehyde	-16.77	-17.53	-18.24		-18.63			
triflic acid	-258.17	-264.27	-265.15		-268.90		-272.37	

We will use these data to predict the heats of formation of the pure components as liquids before they are mixed. Next, consider how one might account for these solvent effects by hydrogen bonding in the liquid. The equilibrium geometries for ensembles of triflic acid are shown in Figure 8 for ensembles numbering 2, 3 and 5 molecules. Notice how the molecules are arranged in different equilibrium geometries as the number in the ensemble increases. We conjecture that the properties of an ensemble of an infinite size (*i. e.*, the liquid) can be estimated from the properties of a much smaller size by a proper extrapolation. Consider, a plot of the data from Table 4 which examines the behavior of the inverse heat of mixing versus the inverse of the ensemble size, 1/n. A liquid would be modeled by $n \rightarrow \infty$ or $1/n \rightarrow 0$, so that the heat of mixing for the liquid would be the value of enthalpy of the pure component is shown by the intercept (Figure 9). From these data, we estimate that the enthalpy of pure, liquid tolualdehyde is 19.49 kcal/mol and that of pure, liquid triflic acid is -269 kcal/mol; whereas liquid water is predicted to have $\Delta H^{\circ}_{f} = -67.3$ kcal/mol.

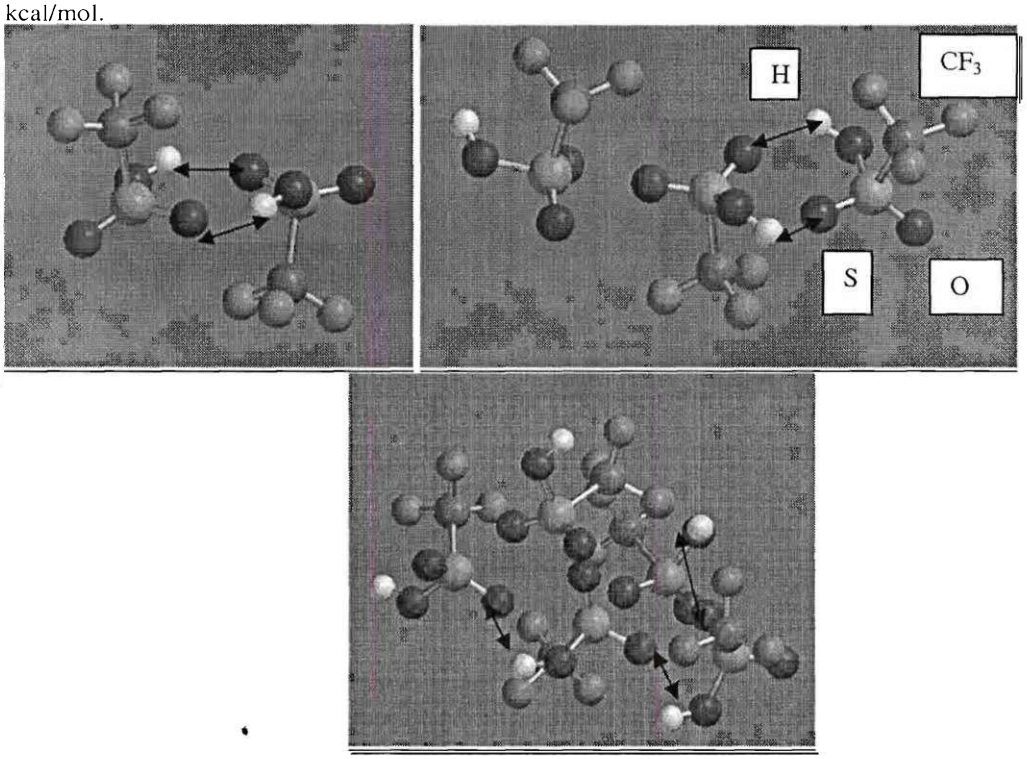


Figure 8 Molecular models of 2, 3, and 5 triflic acid molecules.
Hydrogen bonding showed by arrows

We can compare these predictions with literature data for the thermodynamic properties of pure compounds. For liquid water, the observed heat of formation is -68.32 kcal/mol; whereas, we predicted it to be -67.51 kcal/mol by extrapolating from calculated heats of formation for ensembles numbering 2, 3, 4, 5 8, and 12 molecules. If we include ensembles having only 4, 8, and 12 water molecules, the estimated $\Delta H_{f}^{\circ} = -68.21$ kcal/mol which is 0.11 kcal/mol less than the literature value.

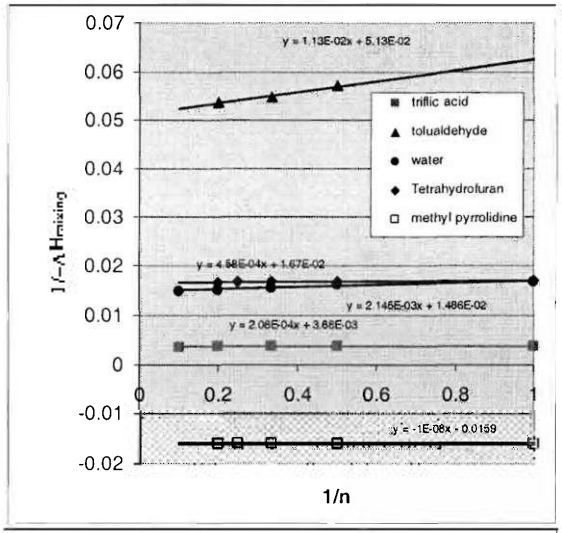


Figure 9 Heats of formation *versus* ensemble number

Now it may be possible to predict the heats of mixing for the water/triflic acid system and for *p*-tolualdehyde/triflic acid system (Figure 10). The predicted heats of mixing were within 1-2 kcal/mol acid of the observed heats of mixing for both water and *p*-tolualdehyde. Some comments are appropriate in describing the method used to estimate the heats of mixing. The semi-empirical quantum mechanics program AM-1 was used to estimate the heats of formation for each of the pure components in the mixture. We used this program to calculate the heat of formation of the pure component sub-ensembles as well as the mixture forming the larger ensemble of this mixture. Then, the heat of mixing was estimated as the difference between the heat of formation for the mixture ensemble and the heats of formation of the sub-ensemble, pure components. As an example, consider a mixture containing m base molecules and n triflic

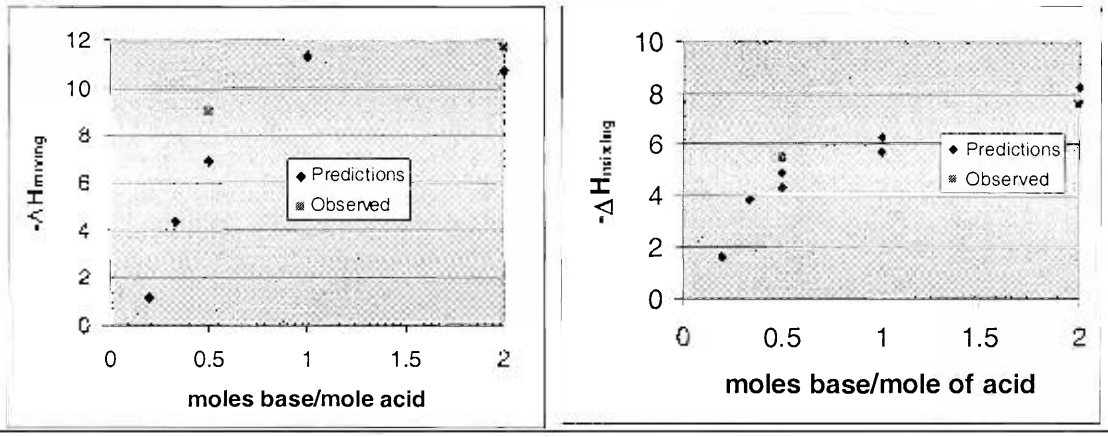


Figure 10 Predicted and observed heats of mixing (kcal/mol acid) for *p*-tolualdehyde (left) & water (right)

acid molecules. The heat of mixing for this ensemble of (n+m) molecules *per mol of acid* is

$$\Delta H^{\circ}_{mixing} = [\Delta H^{\circ}_{mixture}(n + m \text{ molecules}) - n\Delta\hat{H}^{\circ}_{acid}(n \text{ acid molecules}) - m\Delta\hat{H}^{\circ}_{base}(m \text{ base molecules})]/(n)$$

where $\Delta\hat{H}^{\circ}_{i}$ = enthalpy of formation of species i per mole of species i.

Inordinately high heats of mixing result from using the enthalpies of formation for the isolated, pure component. Thus, the present method appears to calculate the observed heats of mixing to an accuracy of 1-2 kcal/mol which is the accuracy of the experimental data (*vide supra*).

We also used this method to predict the heats of mixing for the triflic acid methyl pyrrolidine system (Figure 11). Estimates for the heats of mixing were completed as a function of base/acid ratio (moles/mol) for several different ensemble sizes at a fixed value of base/acid ratio. For example, consider how we modeled the liquid having an acid/base ratio of 1/1 where we used ensembles having either 1 molecule of MPy and one molecule of triflic acid (1/1), 2 molecules of MPy and 2 molecules of acid (3/3), or finally 3 molecules of MPy and 3 molecules of acid (3/3). The estimates of the heat of mixing depended upon the size of the ensemble used to model the system with the largest ensemble showing the highest heat of mixing. This approach was also used to estimate the heat of mixing for base/acid ratios of ½ and 2/1. As before, the highest heat of mixing was estimated from the ensemble having the larger number of molecules. We have placed a trend line through the predictions without those values calculated for the following ensembles: (1/5), (1/2), (2/2), (1/1) and (3/3). The estimates fit by the trend line were 1-3

kcal/mol less than the values observed at the same ratio of base/acid. The predicted trend is for increased values for the heats of mixing, kcal/mol acid, for increasing base/acid ratio at a fixed amount of acid. Apparently, the predictions from small ensembles are less accurate than those developed from larger ensembles at any base/acid ratio. This conclusion suggests that the AM1 calculation does not accurately predict interactions well and that some added calculations were necessary to simulate well the solvent effects. In these calculations we simulated solvent effects by recalculating the properties of the fluid using the same base acid ratio but using larger ensembles of the mixture. The weakness of this approach is that the proper ensemble size is not apparent *a priori*.

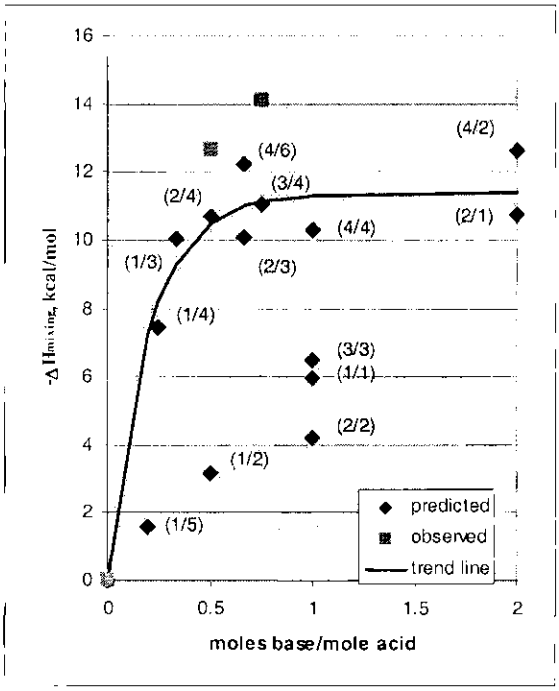


Figure 11 Enthalpies of Mixing for MPy-HOTf

Modeling the drying of triflic acid by MPy. Hommeltoft[45] reported the Haldor-Topsoe process for drying triflic acid using the base, methyl pyrrolidine. At steady state, the wet acid [HOTf n(H_2O)] is contacted with the monoacid salt of methyl pyrrolidine [MPy (HOTf)] to give the intermediate product which is then distilled under vacuum. For the purpose of modeling by quantum mechanics, we let n = 1 as the model for the wet acid. Thus, the process stoichiometry is given by

HOTf-H_2O (s) + HOTf-MPy (liq) → $(HOTf)_2$-H_2O-MPy (liq); ΔH°_1 = 6.80 kcal/mol

$(HOTf)_2$-H_2O-MPy (liq) → H_2O (g) + $(HOTf)_2$-MPy (liq); ΔH°_2 = -6.58 kcal/mol

$(HOTf)_2$-MPy (liq) → HOTf (g) + HOTf-MPy (liq); ΔH°_3 = 11.13 kcal/mol

The enthalpies for each stoichiometric step were estimated from the enthalpies of formation of the species shown in Fig. 12. These species were modeled using the MNDO semi-empirical program to develop the optimized geometries and also to calculate the enthalpies of formation. The modeling predicts that the mixing step is endothermic (6.8 kcal/mol); whereas, the vaporization of water from the intermediate $(HOTf)_2$-H_2O-MPy is exothermic (-6.58 kcal/mol). The last step to produce anhydrous HOTf vapors is endothermic (11.13 kcal/mol).

We have also reproduced this procedure in our laboratory[51] and noticed that a small amount of heat was evolved when the monohydrate of triflic acid was combined with the MPy-HOTf salt. When this liquid was placed under a vacuum of 1 Torr, the system began to "boil" to give vapor water which was collected in a condenser held at dry ice/acetone temperature. After this vaporization was completed, the temperature in the still was increased to 90°C to initiate the vaporization of dry triflic acid at 1 Torr vacuum. The dry HOTf vapors were collected in another flask thermostatted at 0°C and this was continued until the vaporization stopped. At this point the still contained the salt. HOTf-MPy. The MNDO model appears to predict, qualitatively, the heat effects that we observe during the drying process using methyl pyrrolidine. We repeated this modeling using the MNDO-optimized geometries of all species but used the semi-empirical program AM-1 to estimate the single-point-enthalpies of formation. We believe that the actual enthalpies for each step were bound by these two sets of estimates. It is appropriate to comment on the equilibrium geometries determined for this set of species. The MNDO program gave an intermediate, $(HOTf)_2$-H_2O-MPy, with a distance between the oxygen atom in the water molecule and the triflic acid proton that was 2.755 Å whereas the other two models (AM-1 and PM3) gave equilibrium geometries where this distance was 2.018 and 1.758 Å, respectively. Thus, we believe that the model of this intermediate developed from the MNDO program is consistent with our observations that water is removed from the liquid by a room temperature evacuation treatment in that a bond length of 2.755 Å for H_2O—

HOTf suggests a weak, hydrogen bonding interaction as might be observed in liquid water. Subsequently, we modeled ensembles of water containing 4, 6, and 10 molecules using MNDO. The equilibrium geometries showed H_2O—H_2O bond distances of 3.2 to 3.4 Å. This additional modeling of water ensembles suggest to us that the interactions between H_2O and HOTf in the intermediate, $(HOTf)_2$-H_2O-MPy, are similar to those observed between water molecules in "liquid water". From such a weak bond, one might expect that the water molecules could be removed by evacuation at room temperature, which is what we observed.

Predictions—Enthalpy of mixing for ternary mixtures. We reported the heats of mixing when mixtures containing HOTf and a base were combined: 1) HOTf/*p*-tolauladehyde mixed with water ($-\Delta H_{mixing}$ = -34.5 kJ/mol acid) and 2) HOTf/MPy mixed with HOTf/water ($-\Delta H_{mixing}$ = -5.8 kJ/mol acid). Now we attempt to model the heat effects using AM1. Consider first the modeling of the first combination; where we estimate the enthalpy of formation for the combination of 1 *p*-toluladehyde molecule with 2 triflic acid molecules so as to estimate the enthalpy of formation of a liquid containing 1 g-mol of *p*-toluladehyde and 2 g-moles of HOTf. The AM1 program showed an enthalpy of formation = -553.91 kcal/mol. Next, the same software was used to estimate the enthalpy of formation for a mixture containing 4 g-mole of water using as a model 4 molecules of water ($\Delta H^\circ_{formation}$ = -260.895 kcal/mol). Finally, the combined mixture was modeled as 2 molecules of HOTf, 1 molecule of *p*-toluadehyde, and 4 molecules of water having an enthalpy of formation = -821.8 kcal/mol. The heat of mixing was estimated to be (-821.8 + 260.9 + 553.9)kcal/2 mol HOTf = -7 kcal/2 mol acid = -3.5 kcal/mol acid. The observed heat effect was -34.5 kJ/mol acid = -8.2 kcal/mol acid. Thus, the AM1 software underestimated the heat of mixing by a factor of 2. In subsequent attempts to model the same system, we used larger ensemble sizes (Table 5).

Table 5 Predictions of ΔH°_f for ternary system: *p*-tolualdehyde/triflic acid/water

Species	$-\Delta H^\circ_{f,predicted}$	kcal/mol acid		
	n=1	2	3	∞
$[p\text{-tal}]_n[HOTf]_{2n}$	276.96	280.1	-280.0	281.9
	m=4	8	12	∞
$[H_2O]_m$	65.22	66.47	67.34	68.21
	n=1/m=4	2/8	3/12	∞
$[p\text{-tal}]_n[HOTf]_{2n}[H_2O]_m$	136.97	141.27	140.64	143.51
$-\Delta H^\circ_{mixing,\ predicted}$	3.50	10.81	7.27	12.19

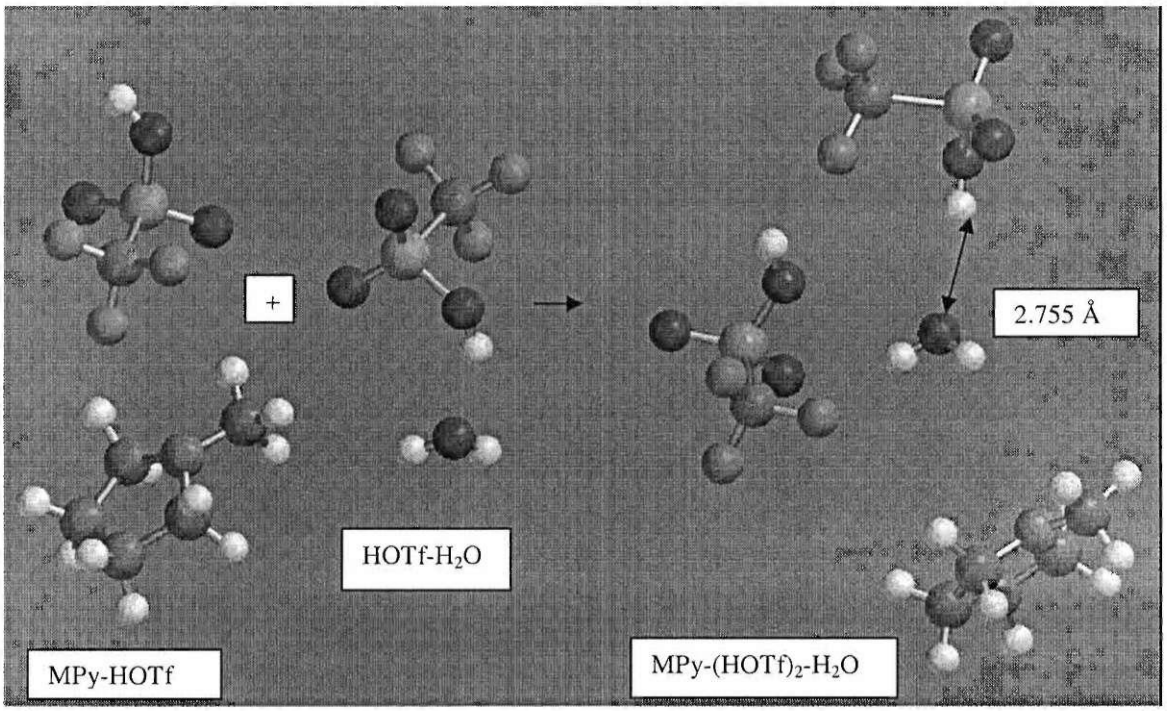


Figure 12 MPy-HOTf salt is added to the monohydrate of triflic acid to yield MPy-$(HOTf)_2$-water. Modeled by the MNDO semi-empirical program.

The actual enthalpy of mixing for this system was -34.5 kJ/mol acid = -8.2 kcal/mol acid. Notice that our estimates for this heat of mixing depended upon the ensemble size and these values range from -3.5 kcal/mol for the smallest ensemble size up to a value of -10.8 kcal/mol for the ensemble containing 2 *p*-tal molecules, 4 HOTf molecules and 8 water molecules. The value for the enthalpy of mixing was -12.19 kcal/mole acid derived from an extrapolation of these estimates to predict the properties of an ensemble of infinite size, which is 4 kcal/mol acid greater than the observed value. On the other hand, we notice that the values of the enthalpies of formation for the mixture of *p*-tal/HOTf actually decreased slightly (0.1 kcal/mol acid) for increasing the ensemble size from n = 2 to n = 3. The same comment can be made for the resulting mixture of *p*-tal/HOTf/water where the enthalpy of formation decreased from -141.3 to -140.6 kcal/mole acid when the ensemble size increased from (n=2)/(m=8) to (n=3)/(m=12). These trends suggest to us an alternative way to extrapolate the data in which we average the values of the heat of formation for the largest two ensemble size. By this method, we estimate a heat of mixing of -6.4 kcal/mol acid which is only 1.8 kcal/mol different from the observed heat of mixing. While none of these estimates correctly predicted the observed heat of mixing, the estimates were surprising close to the observed values given the low level modeling used in estimating the enthalpies of formation.

We attempted to use this method to model the combination of HOTf/MPy with HOTf/water. The mixture containing 1.5 g-mol of HOTf and 1 g-mol of MPy was modeled by 3 molecules of HOTf and 2 molecules of water. The resulting enthalpy of formation was -817.59 kcal/mol. The HOTf/water mixture was modeled by using 1 molecule of HOTf with 2 molecules of water to show an enthalpy of formation = -392.392 kcal/mol. This value will be doubled when used to determine the enthalpy of mixing because the model for HOTf/MPy used twice the number of moles of reagent to create model (*i. e.*, 3 molecules of HOTf and 2 molecules of MPy to model a mixture containing 1.5 g-mol of HOTf and 1 g-mol of MPy). The resultant liquid mixture was modeled using 5 molecules of HOTf, 2 molecules of MPy, and 4 molecules of water to show an enthalpy of formation = -1622.4 kcal/mol. The heat of mixing was determined as follows: [-1622.4 + 817.59 + 2(392.392)] kcal/5 mol acid = -20.03 kcal/5 moles acid = -4 kcal/mole of acid. The observed heat effect was -5.8 kJ/mol acid = -1.4 kcal/mol acid. In this case, the AM1 model overpredicted the heat of mixing by a factor of 3. We did not attempt to predict the properties for larger multiples of this system as the number of electrons was very large even for these ensembles.

Predictions--Reactions. We attempted to model the carbonylation of substrate in super acid using a method that proved fruitful for an isomerization.[49] By this method, we simulated the isomerization reaction by a sequence of reaction followed by equilibration of carbocations. These results showed that toluene can be protonated in super acid solution, and that other methyl-substituted arenes will isomerize in super acids but arenes substituted with electron withdrawing groups (*e. g.*, aldehyde, ester) were not isomerized. Moreover, we were able to predict well the regioselectivity of the isomerizations by assuming that the free energies of the carbocations determined the product distribution.

The direct interaction between CO and toluene is thermodynamically unfavored, but there are two different possible pathways that lead to the insertion of CO into the aromatic C-H bond, when either the arene or the CO is protonated. The conventional mechanism of the Gattermann-Koch formulation in superacidic media have been extensively investigated by Olah *et al.* [46] and other groups, [47] and the proposed reaction mechanism is illustrated in eq. 8-9:

$$CO + H^+ \leftrightarrows HCO^+ \quad \text{Eq. (8)}$$

$$HCO^+ + ArH \leftrightarrows [ArH{\cdot}HCO^+] \rightarrow [ArHC(H)O^+] \rightarrow ArCOH + H^+ \quad \text{Eq. (9)}$$

$$\pi\text{-complex} \qquad \sigma\text{-complex}$$

In this reaction pathway the formation and the attack of the electrophile are separate steps. The diprotonated, protoformyl dication $HCOH^{2+}$ was also suggested as the active species.[52]

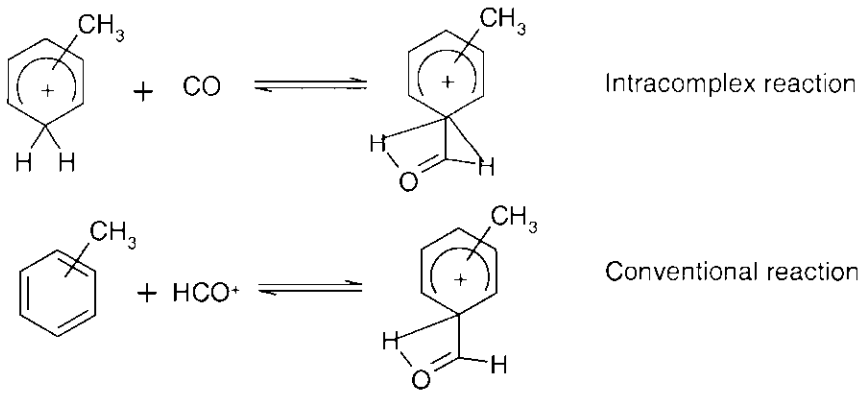
An alternative mechanism, the intracomplex reaction, suggests that the formation of the electrophile HCO^+ occurs within the complex of proelectrophile and protonated aromatic compounds,[53] eq. 10-11:

$$ArH + H^+ \leftrightarrows ArH_2^+ \quad \text{Eq. (10)}$$

$$\underset{\pi\text{-complex}}{ArH_2^+ + CO} \rightarrow \underset{\pi\text{-complex}}{[ArH_2^+CO]} \leftrightarrows [ArH^+C(H)O^+] \rightarrow \underset{\sigma\text{-complex}}{[ArHC(H)O^+]} \rightarrow ArCOH + H^+ \quad \text{Eq. (11)}$$

The mechanisms differ by the way they orchestrate the regioselectivity of the reaction. In the case of the intracomplex reaction, the regioselectivity can be heavily influenced by the stability of the protonated arene isomer. The logical extension of this concept, is that certain regioisomers of the products are selectively excluded from the product slate when the free energy of the corresponding protonated arene isomer is much higher than the free energies of the other, protonated arene isomers.

We attempted to understand the selectivity of the process by modeling a pseudo-transition state of this reaction, which is a transition state that corresponds to the formation of a complex (eq. 8-9 and 10-11) from already formed reactant cationic species (ArH_2^+ or HCO^+). This approach allows us a closer look the species that can be accountable for the regioselectivity of the formylation reaction. The geometries of these states (scheme 1) were obtained by the built-in module in Titan software.



The results are summarized in the table 6.

Table 6 Estimates of transition state energies

Reactants	$E^{\#}$, kcal/mol
o-TolH_2^+ + CO	**71.6**
m-TolH_2^+ + CO	-5.9
p-TolH_2^+ + CO	-3.4
o-Tol + HCO^+	-51.5
m-Tol + HCO^+	-52.7
p-Tol + HCO^+	-53.1

These predictions suggest that the formation of pseudo-transition state, starting with the protonated arene and CO, is thermodynamically favorable in all cases except with protonated *o*-toluene, where the barrier is too high (71.6 kcal/mol). On the other hand, all transition states appear equally favored when starting with arene and protonated CO. Consequently, it seems likely that *o*-tolualdehyde is formed by the conventional mechanism. These predictions also show very little discrimination among the isomers of tolualdehyde by the conventional process (protonated CO). It is also evident that high regioselectivity of the carbonylation process cannot be explained only by consideration of that pseudo-transition state formation since both the *m*- and *p*-isomer are equally favored by the transition state originating from the intracomplex reaction.

In that respect we examined other species that in theory may isomerize in super acid and affect the tolualdehyde isomers distribution. Notice that in the conventional reaction pathway (eq. 8-9) the only species that may determine the distribution of the final product is the protonated tolualdehydes and their

free energy will determine the product slate. When taking into the consideration intracomplex reaction mechanism (eq. 10-11), in addition the free energy of protonated toluene isomers must be accounted for. These two reaction schemes involve a type of conversion process in which the relative energies among the regioisomers of the carbocations are the governing consideration. Such regio comparisons belong to a class of reactions known as isodesmic reactions (*i. e.,* conformation changes). It has been noted that Hartree-Fock, H-F, models do a better job of describing these reactions than the electron correlation models (DFT and LMP2).

We attempted to use this strategy here to predict the carbonylation of toluene in super acid. The results are shown in Table 6. The results are listed for six methods of prediction ranging from the approximate AM-1, semi-empirical method; the Hartree-Fock method using the 6-31G** basis set for the orbitals; the B3LYP/6-31G**; LMP2/6-31G**; and LMP2/cc-pVTZ (for the later the heat of formation was additionally calculated by SPE methods using resulted equilibrium geometry configuration). The first set of calculations, in columns 2-4, assume that the protonation of the toluene is the carbocation equilibrium that defines the product distribution (scheme 1); whereas the second set of calculations, columns 5-7, assumes that the protonated tolualdehyde equilibrium sets the product distribution (scheme 2). The actual results for this carbonylation were *o/m/p* = 5.5/0.5/94%. The predictions of products' isomer distribution based upon the protonated tolualdehyde equilibrium (scheme 2) do not agree with the observed results in that the *meta*-isomer appears in amounts from 1-9% with only small amounts of the *o*-isomer. If we turn to the results predicted for the protonated toluene as the defining equilibrium, then the Hartree-Fock and the SPE calculations appear to model well the results that we observed. All of the models confirm that the scheme involving the protonation of the tolualdehyde isomers is not an acceptable description of the reaction pathway leading to the actual isomers distribution. This conclusion is also confirmed by the experimental result that *p*-tolualdehyde cannot be isomerized in neat triflic acid at room temperature.[54]

Thus, we are left with eq. 10-11 as one acceptable mechanism for predicting the regioselectivity of the carbonylation. These predictions show that the most stable carbocation is the *para*-isomer followed by the ortho. The meta-isomer is predicted to be present in minor amounts by all of the models. Thus, one might expect that very little *m*-tolualdehyde would be formed if the mechanism involved the intracomplex reaction. On the other hand, if the classical mechanism were operative, then the protonated CO should attack the ring of the neutral arene at any position which would lead to a protonated aldehyde complex. The stabilities of these protonated aldehyde complexes, shown in Table 7, suggest that some *m*-tolualdehyde should come out of the transition state.

Table 7 Predicted product distributions for carbonylation of toluene in super acid

Prediction	protonated toluene			protonated tolualdehyde		
Methods	*% ortho*	*% meta*	*% para*	*% ortho*	*% meta*	*% para*
AM-1	25.53	0.10	74.38	4.11	4.24	91.65
HF/6-31G**	5.83	0.00	94.17	0.49	1.22	98.29
B3LYP/6-31G**	17.85	0.04	82.15	0.51	2.98	96.51
LMP2/6-31G**	15.83	0.10	84.07	0.79	8.30	90.40
LMP2/cc-pVTZ	12.27	0.07	87.67	0.70	6.05	93.25
SPE ΔH_f	6.46	0.04	93.50	1.55	2.26	96.19
Experimental	5.5	0.5	94.0			

We reported evidence[55] from ^{1}H-NMR and mass spectrometry that toluene exchanged D for H when perdeuterated toluene was mixed with the proton-form of triflic acid at room temperature after 2 h of mixing; however, no D-H exchange was observed between perdeuterated toluene and the proton-form of a weaker acid: trifluoroacetic acid. These H-D exchange results confirm that the protonated carbocations of toluene were formed in the presence of a strong acid such as triflic acid. Thus, we now examine the predictions based on eq. 1 and find that the H-F and SPE methods predict well the regioselectivity of the carbonylation reaction. This result is not unexpected since the equilibration reaction between protonated, toluene isomers is an isodesmic reaction that should be predicted well by the H-F method.

Discussion

Physical properties. Toluene is only partially soluble in triflic acid (Figure 1, solubility limit = 27 mol%) and therefore large excesses of triflic acid is required to completely solubilize toluene. This partial solubility of toluene in triflic acid would be of importance for the commercial practice of this science. We found that the toluene solubility in the acid could be improved if a small amount of the product, tolualdehyde, was used as a co-solvent. When *p*-tolualdehyde was added in as little as 12 mol% in the triflic acid, the toluene became miscible (Figure 1). While this approach may seem an appealing alternative to using large excesses of triflic acid, tolualdehyde is a "poison" to the triflic acid in that it forms an adduct with the acid. The carbonylation reaction is completely inhibited when *p*-tolualdehyde is added to the acid in as little as 10 mol % (Figure 6).

When the reaction has gone to completion, water is added to 1) stop the reaction and 2) assist in separating the product from the triflic acid. The minimum amount of water needed to free the aldehyde from the acid was 13 mol of water per mol of acid initially present in the product mixture. This very large amount of water needed to free the product presents a real challenge to the commercial implementation of this technology since the reaction mixture also contains large excesses of triflic acid (~10 mol acid/mol substrate) in order to produce high conversions. This high requirement for water to "break" the acid/aldehyde complex can be understood better when one considers the phase diagram for forming the triflic acid/hydrates.[56] Water forms a sequence of hydrates with triflic acid (HOTf nH_2O; n = ½, 1, 2, 4, and 5) at room temperature which also forms solutions with excess water. Some of these hydrates are solid at room temperature (*e. g.*, hemihydrate, hydrate, *etc*). The resulting "wet" triflic acid is unreactive for further carbonylation reaction if the mole fraction of water is greater than 1/3 (Figure 6). The "wet" acid can be regenerated by removing the water either by vacuum distillation[57] or with the use of a tertiary amine drying agent.[45] The patented process for drying triflic acid involves the contacting of the "wet" acid with the complex of methyl pyrrolidine and HOTf to yield another complex of the acid and the dihydrate of the methyl pyrrolidine. This complex is then heated to decompose it leaving water vapor and a liquid containing the acid and methyl pyrrolidine. This liquid is heated further to yield a dry triflic acid vapor and the HOTf/MPy complex which may be recycled to dry more "wet" acid. These processes all show excess heats of mixing and thus, we reported the enthalpy changes which attended the mixing of certain binary and ternary mixtures so that estimates of heat loads on the commercial mixing devices could be developed for design considerations.

Heats of mixing were estimated from the temperature rise in a batch calorimeter of a simple design constructed from Pyrex™ glass. It was of interest to us to know the heat liberated upon adding water to pure triflic acid as a means of estimating the amount of heat that might be produced upon adding water to the product mixture. We estimated that 40-50 kJ (~10 kcal) of heat were liberated for every mole of acid present when as little as 1-2 moles of water is added. When *p*-tolualdehyde was added to pure triflic acid, the heat of mixing was lower than that found for water (23-32 kJ/mol acid, ~ 7 kcal/mol acid). Thus, one might expect that water could displace the aldehyde from triflic acid and thus free it. From the thermal chemistry alone, our results also suggest that the heats of mixing for methyl pyrrolidine, tetrahydrofuran (THF), and acetonitrile with the acid were all very similar (~45 kJ/mol acid). Further work will be needed to show that THF and acetonitrile can function as drying agents for wet triflic acid.

It was of considerable interest to us to model the heat effects upon mixing using simple, semi-empirical quantum mechanics programs. The approach is straight forward in that the quantum mechanics software calculates the enthalpy of formation of the species at 298 K. However, the software completes these calculations for the "isolated" species meaning that solvent effects must be modeled by proper choice of the ensemble size. Since the heat of mixing arises largely from molecular interactions between the species, the modeler must consider a proper way to represent the liquid mixtures. We chose to calculate the enthalpy for ensembles of different sizes and then to extrapolate these results to predict the enthalpy of formation for an ensemble of infinite size so as to model the properties of the desired species as a liquid. As a test case, we modeled the properties of water molecules in ensembles ranging from 1-12 molecules (Table 4). The enthalpy of formation for the isolated water molecule (n = 1, ΔH°_f = -59.24 kcal/mol) was nearly the same as that for the ΔH°_f for gaseous water (ΔH°_f = -57.8 kcal/mol). The predicted values of the

ΔH°_{f} became larger in a systematic fashion as the ensemble size increased. One method to extrapolate these data to predict the properties of an ensemble of infinite size is to use a reciprocal plot, where the inverse of the ΔH°_{f} is plotted versus the inverse of the ensemble size, n. The asymptotic value of ΔH°_{f} is the value of the intercept. This extrapolated value for the heat of formation for water was -67.3 kcal/mol when all of the predicted values were used and -68.2 kcal/mol when we used only the predictions from the larger ensembles. The literature value for liquid water is 68.32 kcal/mol. While both estimates were close to the literature value, the weakness of this approach is in choosing the proper subset of ensembles to model the real system. While our results suggest that the predictions from larger ensembles seem to produce better estimates of the heat of formation for the liquid, it may not be possible to use very large ensembles due to the limitation of computing power. Thus, the problem to be solved is choosing the proper size of model for the real system that can be determined from a model of suitable size to compute the energies. We used this approach to predict the heats of mixing that we reported earlier in this manuscript.

For each case we used the smallest ensemble sizes to predict the heats of mixing for triflic acid with each of the following bases to form the binary mixture: 1) *p*-tolualdehyde and 2) water. We were able to predict the heat of mixing to values within 1-2 kcal/mol acid of the observed values (Figure 10). For these two systems, it appeared that successful predictions of the enthalpy of mixing could be achieved with the smaller ensembles as models for the liquids. However, such was not the case for modeling the triflic acid/methyl pyrrolidine system (Figure 11). The enthalpies of mixing predicted from using models of smaller ensembles were much lower than the corresponding values predicted from the larger ensembles. Close agreement were found between predicted and observed heats of mixing for the larger ensembles only. Again, we can find ensembles that model well the observed heat effect, but it is not obvious how we should make this choice of ensemble size *a priori*. As such, these modeling efforts are better characterized as attempts to correlate data.

Given the difficulty in choosing the proper model for these binary mixtures system, our efforts to model the ternary mixtures were only modestly successful. Consider, the ternary system of triflic acid/*p*-tolualdehyde/water. This system mimics the neutralization process whereby the aldehyde/acid adduct is broken with the addition of water. Modeling the actual reaction mixture is beyond the capabilities of the software in that too many molecules would be required: 1 *p*-tolualdehyde, 10 triflic acid, and 130 water molecules. Therefore, we compromised with a system containing the reagents in the following proportions: 1 mol *p*-tolualdehyde; 2 mol triflic acid; and 4 mol of water. We also had collected thermal data for this system when a mixture of *p*-tolualdehyde/HOTf (1/2 mol) was combined with water (4 mol) to show a heat effect of ~ -35 kJ/mol acid. Using models constructed from the smallest ensembles (1 *p*-tolualdehyde, 2 HOTf, and 4 water molecules), we estimated the heat of mixing to be ~ ½ that of the observed heat of mixing (-4 kcal/mol acid *versus* -8 kcal/mol acid, observed). We repeated these efforts with models containing 1) 2 *p*-tolualdehyde/4 HOTf/8 water molecules and 2) 3 *p*-tolualdehyde/6 HOTf/12 water molecules.

Chemical properties. The evidence for direct carbonylation of toluene by super acids is compelling even though the reaction is not favored by the Gibbs free energy of reaction based upon the free energies of the neutral reactants and product. We estimated this free energy of reaction to be +8 kcal/mol when the reaction is completed at 25°C.[58] This slightly positive, free energy of reaction suggests that the reaction cannot go to completion if the neutral species are equilibrated. Therefore, we are left to conclude, along with others,[46, 59, 60] that the reactants and products are not the only species present and that some other species may be present that serve to lower the free energy for its formation. The work-up precludes the neutral tolualdehydes from reacting backwards through the equilibrium because the super acid is converted to a weaker acid by the addition of water to the reaction mixture. Thus, when considering the equilibrium, one must include protonated species. This macroscopic thermodynamic calculation relates to the microscopic details of the mechanism insofar that protonated species are certainly present the mechanism. For example, it is known that CO is thermodynamically-forbidden from direction insertion into an aromatic C-H bond; however, this restriction may not be operating when the species are protonated, either the CO or the arene.

The carbonylation of arenes in large excess of triflic acid appears to be a fruitful path to the synthesis of arene aldehydes without the formation of the subsequent reaction products of the substrate with

the aldehyde. We speculated that the large excess of the triflic acid reduced the rate of the bimolecular reaction of the aldehyde with the substrate as a result of a solvent dilution effect.[54] The products observed from the toluene carbonylation, *o*- and *p*-tolualdehyde, are expected from a consideration of the published reaction mechanisms.[46] The carbonylation of *p*-xylene produced two major products: 2,5-dimethylbenzaldehyde (97%) and 2, 4-dimethylbenzaldehyde (3%). We did not detect any of the other three possible regioisomers (2,3-; 3,4-; & 3,5-dimethylbenzaldehyde). The 2,5-dimethylbenzaldehyde is the expected product from direct carbonylation of *p*-xylene. The product 2,4-dimethylbenzaldehyde could occur as the result of isomerization of the substrate followed by carbonylation of the *m*-xylene intermediate. This isomerization was reported to occur at 80°C when the super acid mixture HF/BF_3 was contacted with a xylene to produce exclusively *m*-xylene.[47] We attempted to isomerize *p*-xylene in triflic acid at room temperature without much success. Thus, we explain the results of *p*-xylene carbonylation by a model in which isomerization and carbonylation of the substrate are in competition, for which the kinetics of each reaction directly influenced by the reaction temperature. At room temperature, the carbonylation is much faster than the isomerization, whereas the converse is true at 80°C. The slightly lower reactivity of the *p*-xylene over the toluene carbonylation is reasonable in the light of our earlier results where *p*-xylene slowed the reaction rate of acid-catalyzed isomerization of dimethylbiphenyl when it was used as a solvent. We reasoned that *p*-xylene was also easily protonated by the triflic acid and therefore, the solvent *p*-xylene competed with the substrate, dimethylbiphenyl, for the triflic acid.

Carbonylation of mesitylene at room temperature showed one major product (2,4,6-trimethylbenzaldehyde). The conversion of the substrate was 71% These results suggest that the substrate did not isomerize at the reaction conditions before the carbonylation and this suggestion is consistent with the data of McCaulay and Lien[47] who showed that mesitylene was the thermodynamically favored product in super acid media when the temperature was 100°C. Thus, one would expect no rearrangement of the mesitylene substrate. The lower reactivity of mesitylene when compared to *p*-xylene and toluene is consistent with a mechanism whereby intracomplexes are formed by the arene and the triflic acid. The stability of these complexes with the triflic acid obeys the order:

$$\text{mesitylene-H}^+ > \textit{p}\text{-xylene-H}^+ > \text{toluene-H}^+$$

The most stable carbocation is the least reactive towards reaction with the CO; thus, one predicts the following order of reactivity towards arene carbonylation: toluene > *p*-xylene > mesitylene. In the case of 1,2,4-trimethylbenzene, we expect that its stability towards forming a carbocation will be less than that demonstrated by mesitylene since stabilization of the charge is more favorable in mesitylene.

The kinetics of the toluene carbonylation cannot be explained completely by a simple first order treatment. The product aldehyde may react further with the substrate to make the ditolymethane when the acid/substrate ratio is low. This second order reaction can be minimized by adding triflic acid as a solvent. As the tolualdehyde product accumulates, the rate to form ditolymethane increases as part of the reaction sequence:

$$\text{toluene} + \text{CO} \rightarrow \text{tolualdehyde}$$
$$\text{tolualdehyde} + \text{toluene} \rightarrow \text{ditolymethane} + H_2O$$

With the formation of ditolymethane, the reaction rate of toluene conversion decreases as the triflic acid is diluted with water. We showed earlier, that even small amounts of water (~1 wt%) in triflic acid, drastically inhibited the reaction rate.[50] Thus, we expect that the conversion rate should decrease with increasing conversion faster than that predicted by simple first order kinetics.

Supported triflic acid. Efforts were reported to support triflic acid on silica resulting in a "heterogenized" triflic acid.[61] These researchers showed results to suggest that a thermally stable material could be developed by treating silica gel with triflic acid at 150°C in a closed vessel using Freon 113 as the solvent. The resulting solid was active for alkylation of isobutene with *n*-butenes. Attempts were made to demonstrate that the triflic acid could not be leached from the solid. If this material is truly triflic acid grafted onto the surface as the authors claim, then it might be an active carbonylation catalyst. If on the other hand, the catalyst is highly dispersed hemi- and monohydrates of triflic acid, then it will not be active.

Perfluoroalkanesulfonic acids—Carbonylation Agents

[a]"The longer perfluoroalkanesulfonic acids are hydroscopic oily liquids. Distillation of the acid from a mixture of its salt and sulfuric acid gives a hydrated mixture with melting points above 100°C. These acids show the same general solubilities as trifluoromethanesulfonic acid, but are insoluble in benzene, heptane, carbon tetrachloride, and perfluorinated liquids. All of the higher perfluoroalkanesulfonic acids have been prepared by electrochemical fluorination[26]."

"The longer-chain acids and their salts, particularly $C_8F_{17}SO_3H$ and higher, are surface-active agents in aqueous media. They reduce the surface tension of water to levels not possible with hydrocarbon surfactants. The surfactant $C_8F_{17}SO_2N(C_2H_5)CH_2COOK$ [2991-51-7] lowers the surface tension of water to at 0.2 weight percent[62] and exhibits outstanding thermal and chemical stabilities. The potassium salt of perfluorooctanesulfonic acid [2795-39-3], $C_8F_{17}SO_3K$, or perfluoroethylcyclohexanesulfonic acid [335-24-0], C_2F_5-cyclo-(C_6F_{10})-SO_3K , can form a stable foam in hostile media such as chromium trioxide and sulfuric acid where conventional hydrocarbon and silicone surfactants would be destroyed[63]. The ability of these materials to foam concentrated sulfuric acid is utilized to prevent sulfuric acid from aerosoling into the air in industrial situations where chrome plating is done. Instead of forming an aerosol, the sulfuric acid forms a foam blanket on top of the plating bath. The foam derived from $C_8F_{17}SO_3K$ is generally more stable and dense than that derived from C_2F_5-cyclo-(C_6F_{10})-SO_3K . As a result these fluorochemical surfactants are often used in combination to produce desired wetting and foaming activity."

"Generally, derivatives of the longer-chain perfluoroalkanesulfonic acids have a number of unique surface-active properties and have formed a basis for a number of commercial products. Derivatives of *N*-alkyl perfluorooctanesulfonamidoethanol, $C_8F_{17}SO_2N(R)CH_2CH_2OH$, and polymers of *N*-alkyl perfluorooctanesulfonamidoethyl methacrylate, $C_8F_{17}SO_2N(R)CH_2CH_2OCOC(CH_3)\ CH_2$, impart soil, oil, and water repellency to treated fabrics and paper; this forms the basis for 3M's Scotchguard and Scotchban products[64]. Polymers of *N*-alkyl perfluorooctanesulfonamidoethyl acrylates, $C_8F_{17}SO_2N(R)CH_2CH_2OCOC(R')\ CH_2$, with certain hydrocarbon acrylates and methacrylates have also been found to be surface-active agents in organic solvents and water. These polymers have applications in the areas of secondary crude oil recovery and wetting, leveling, and flow control agents[65,66]."

"Higher perfluoroalkanesulfonates are slightly more reactive than triflates toward nucleophilic displacements. The rate constants for acetolysis of methyl nonafluorobutanesulfonate [6401-03-2], methyl trifluoromethanesulfonate [333-27-7], and methyl toluenesulfonate [80-48-8] are 1.49×10^{-4}, 7.13×10^{-5}, and, $3.1 \times 10^{-9}\ s^{-1}$ respectively. This means that the relative reactivities for nonafluorobutanesulfonate, trifluoromethanesulfonate, and toluenesulfonate are 48,000/22,900/1[67]."

Exxon researchers[68,69,70] report in the patent literature the promising use of $CF_3(CF_2)_nSO_3H$ (n = 1, 5, & 7) and $CF_3(CF_2)$-O-$(CF_2)_nSO_3H$ (n = 2) for the carbonylation of arenes. These homologues of triflic acid appear promising for commercial application since it is claimed that the aldehyde products can be separated from the acid catalysts by vaporization. The acid strengths of the perfluoroalkanesulfonic acids are lower than triflic acid and therefore, we expect the yields/conversions of arene carbonylation will be less with the higher molecular weight homologues.[71] Yields are low (1-9%) as a result of low arene conversions while good *para*-selectivity was realized (~93%) with the remaining isomer claimed to be the *ortho* isomer. The proposed process for recovering the perfluoroalkanesulfonic acid uses trifluoroacetic acid as a solvent which assists in breaking the acid/aldehyde complex (*vide infra*) in a separator before sending the products to a wiped film evaporator under a partial vacuum (100°C at 0.25 Torr) where the tolualdehyde was recovered upon heating the fluid. The trifluoroacetic acid solvent was recovered separately from the perfluoroalkanesulfonic acid using a wiped film evaporator. We know that trifluoroacetic acid alone is not sufficiently acidic to catalyze the toluene carbonylation. If the claims of this patent can be repeated in the laboratory, then, it may have warrant further investigation.

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[a] Text in quotes were taken *verbatim* from the report by Patricia Savu, "Fluorine-Containing Polymers, Perfluoroalkanesulfonic Acids", *Kirk-Othmer Encyclopedia of Chemical Technology* (1994), John Wiley & Sons, Inc.

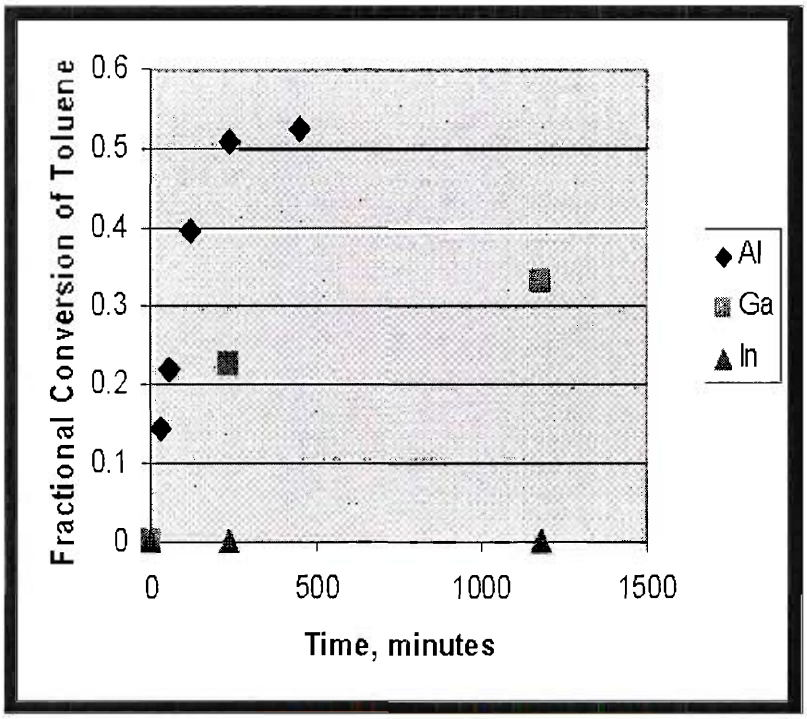


Figure 13 Reactivity of Toluene Carbonylation Towards IL's Developed from *n*-Butyl Imidazolium Chloride-2MCl_3: M = Al, Ga, In.

Ionic Liquids

Chloroaluminate ionic liquids have been reported as efficient conversion agents for toluene carbonylation when the cation was either dialkylimidazolium chloride[72,73] or N-pyridinium chloride[74]. Recently, Wasserscheid,[75] reported another IL conversion agent which is a combination of a cation, such as N-alkyl-4-methylpyridinium chloride, an anion such as $(CF_3SO_2)_2N$-$AlCl_3$. They found extraordinary solubility of $AlCl_3$ in these new IL's, up to 4.5 mol $AlCl_3$/mol of IL. These results are to be contrasted with the results for other chloroaluminate IL's where the maximum solubility is 2 moles $AlCl_3$/mol IL. All of these agents function as stoichiometric reagents, thus the maximum conversion of arene to aldehyde is determined by the maximum amount of aluminum chloride dissolved in the IL. However, the aluminum chloride must be neutralized with water before the product aldehyde is released and because the aluminum chloride is destroyed, these agents will not be commercially- nor environmentally-viable for large volume processes. A breakthrough in this technology is required that will permit 1) the product to be released from the IL, and 2) regeneration of the IL so it can be reused.

A world patent publication described an IL developed from dialkylimidazoium chloride and $InCl_3$, which could be regenerated by heating in vacuum.[76] Although this IL was active for a Freidel-Crafts reaction (toluene + benzoic anhydride) and could be reused several times, we did not observe any products of toluene carbonylation. It appeared prudent to examine an IL derived from a group III-A metal chloride of Lewis acidity intermediate to Al and In: Ga. We show here, Fig. 13, the results of toluene carbonylation with three IL's derived from *n*-butyl-methyl imidazolium chloride and MCl_3, M = Al, Ga, or In. This figure shows that the most active conversion agent for toluene carbonylation is the IL derived from Al, and that Ga shows an activity that was intermediate to those activities of Al and In. Angueira[77] measured the Hammett acidity functions of these three IL's and reported them to be as follows: -13.99 for Al, -13.72 for Ga, and -4.63 for In. From these estimates of Brønsted acid strength, it is clear why the In IL is not active for toluene carbonylation since most researchers suggest that a Hammett acidity function of –12 and more negative is needed to activate the system for carbonylation. However, we did not determine if the Ga IL could be regenerated; this experiment might be worth doing!

Solid Acids

Zeolites. Solid, heterogeneous acids offer an attractive alternative to liquid, homogeneous acids for the functionalization of arenes in that the products/reactants can be easily separated from the solid acids by physical means. One family forming a subset of solid acids is the synthetic, microcrystalline aluminosilicates or zeolites. The well-defined crystal structures of both natural and synthetic zeolites permit selective hydrocarbon transformations. The selectivity of the zeolites can be improved by deactivations of external acid sites with amines, replacement of the cationic sites by transition metal ions by

ion-exchange, or by modification of the silica–alumina ratio. Some zeolites, such as H-ZSM-5 have treated to develop what is claimed to be super acidic character at high temperatures. Accordingly, Hamilton Pittman, *et al.* report the formylation of benzene using a zeolite (13Y) that was modified to contain a copper compound.[78] This solid was treated to high temperature and hydrogen to partially reduce the cupric ion to copper of a lower oxidation state, perhaps even the cuprous state. The reaction conditions were room temperature, equimolar amounts of CO and HCl were added to a pressure of 500 psig. Afterwards, the gases were vented and the system was treated with a desorbing agent (steam, CO_2, or ammonia) to recover the aldehyde from the solid. The selectivity was reported to be 65% but no conversion of the substrate was reported. The solid was regenerated by heating to 325°C so as to remove the desorbing agent.

Recent work in our laboratories[72] and others[79] to characterize the acidity of zeolites and liquid acids show that the X, Y, and ZSM-5 zeolites demonstrate acid strengths that are less than concentrated sulfuric acid. For example, Angueira reports estimates of the Hammett acidity function for X, Y, and ZSM-5 to be as follows: -3.5, -5.2, and –6.2. The threshold value for the Hammett acidity function for a substance to act as a toluene carbonylation agent is –12. Thus, alone, we expect that these zeolites would not be active conversion agents for toluene carbonylation. Hence, we expect that HCl gas is necessary to provide the super acidic Bronsted sites when it combines with extra framework aluminum species that are generated during the high temperature pretreatment with steam. The inclusion of the Cu species may assist in the adsorption of CO.

Stepanov, *et al.*, observed the Koch reaction (olefin or alcohol + CO + H_2O) in zeolite ZSM-5 at room temperature and low pressure.[80] This reaction normally requires concentrated sulfuric acid and CO at 100 bar at room temperature. These NMR results suggest that the ZSM-5 may show strong acidity.

Modified zeolites. We reported the synthesis and characterization of a microporous zeolite containing small amounts of Si (5 mol%) and much larger amounts of Al and P oxides which were developed from reacting $SiCl_4$ gas with aluminum phosphates having the same structure as ZSM-5 ($AlPO_4$-5).[81] Whereas, the parent $AlPO_4$-5 was unreactive to the probe reaction, toluene-methylation to form mixed xylenes, the modified product solid was significantly more active and gave unusually high yields of *o*-xylene (51%), *p*-xylene (27%) and the balance was *m*-xylene. The regioselectivity expected of a Friedel-Crafts alkylation at these conditions was observed: *o*/*p*/*m* = 54%/29%/17%. We explained these results by assuming that $AlCl_3$ was present in the solid as a result of the reaction of $SiCl_4$ by the following:

$$4SiCl_4 + [Al_2O_3P_2O_5]_n \rightarrow 4SiO_4[Al_2O_3P_2O_5]_{(n-1)} + 2AlCl_3 + 2PCl_5$$

The $AlCl_3$, when combined with small amounts of water, created a super acid that could function as a Friedel-Crafts catalyst. If this agent does show super acidity, this modified $AlPO_4$-5 may be an ideal candidate for further testing as an arene formylation agent. Others report the post-synthesis modification of zeolites to extract aluminum oxides from the framework of aluminosilicates by the action of triflic acid.[82] They found optimum conditions for aluminum extraction that balanced the amount of aluminum extracted *versus* the surface area/pore volume of the modified solids. The enhanced reactivity of these solids towards the MTBE synthesis reaction was explained by the formation of reactive, extra-framework Al species rather than a result of triflic acid deposited on the surface of the zeolites. It remains to be proven than the extra-framework Al can be a toluene carbonylation catalyst. Still others report the activity of zeolites combined with $AlCl_3$ for carbonylating benzene at modest conditions.[83] It appears that the zeolite provides protons to initiate the electrophilic substitution reaction but these researchers did not attempt to reuse the solid.

Modified zirconia's. Sulfated zirconia has been reported by others as a conversion agent for toluene carbonylation although the yields were very low.[84] The best catalyst also contained Fe/Mn or Cr and the solid was reported to accumulate a black residue on the surface which prevented the catalyst from operating continuously. The observed yields of 1-2% corresponded to just about a stoichiometric layer of products on the surface. Additional work reported by Clingenpeel, *et al.* provided fundamental understanding to the mechanism of reaction over these solids.[85] For these agents to be used in a continuous process, some means must be developed to remove the surface coating continuously. Funamoto, *et al.* report a process whereby the isomerization of *n*-butane over sulfated zirconia was maintained at a steady activity when the pressure and temperature were adjusted to show super critical conditions.[86] This catalyst when used under

lower pressure conditions shows the normal decline in activity associated with deposition of coke on the surface. The stable reactivity of this catalyst under supercritical *n*-butane suggests that this supercritical solvent is capable of removing solid residue from the surface that would normally block the reactive sites to further reaction. As such, this approach may be adapted for use with arene formylation provided that we can find a supercritical solvent that can displace the aldehyde from the surface in a separate process step.

Mesoporous zirconia was treated with triflic acid in an attempt to improve its reactivity towards a Brønsted demanding reaction: benzoylation of biphenyl.[87] They claim that the subject catalysts is 6 times more active than sulfated zirconia but its turnover rate was ~ $1/5^{TH}$ that of triflic acid. If this claim can be substantiated, then the modified mesoporous material could be useful as an arene carbonylation catalyst provided that it can be reused.

Acidic polymers. Polyfluorinated sulfonic acids are considerably stronger Brønsted acids than the conventional styrene-based materials ($-H_0$~10). They are also more thermally robust. Successful applications of these strong solid acids include the bimolecular conversion of alcohols to ethers. An interesting development in this context is the preparation of composites of silica with fluorinated sulfonic acid resins (*e.g.*, Nafion). These can be considered as effectively being higher surface area forms of the highly acidic polymer with a commensurate increase in catalytic activity of up to 10^3 and allows catalysis of reactions not normally considered possible for conventional ion exchange resins. Thus a Nafion resin/silica catalyst is active for the alkylation of benzene with C_{12} olefins to give the very large-scale detergent intermediate dodecylbenzene. Its particularly important to note that very little branched alkylates on other by-products are formed. Selectivity in this application is very good although the catalyst is significantly less active in such alkylations than supported aluminium chloride or sulfated zirconia.[88] We confirmed the inactivity of these composite materials towards the toluene carbonylation reaction.[89]

An enhanced activity form of this resin has been reported by Sun, *et al.*, recently.[90] They report that the composite is nearly 100 times the activity of the Nafion beads and even more active than strong triflic acid based on the same amount of acid groups present in the catalysts based upon yields/conversions of the benzylation reaction. This reaction does not demand the acid strength required of the arene formylation reactions so it remains to be proven that the Nafion/silica composite is a viable candidate for arene carbonylation.

Several reports[91,92] in the patent literature remark on the enhanced yields of arene aldehydes obtained over systems that combined liquid triflic acid and either porous organic or porous inorganic solids decorated with noble metals (*e. g.*, Pd). They also report much lower yields obtained from just the solids alone. It appears that these findings can be explained in a simple fashion for which the liquid and solid systems are acting independently and thus no new science or technology can be derived from these patents.

A general comment can be made regarding the use of solid acids as carbonylation agents. To be effective and not limited by equilibrium thermodynamics, acidic sites on the surface must form adducts with the aldehyde product just as was observed in liquid acids and ionic liquids. Therefore, the solid acid will act as a stoichiometric reagent for the formation of the aldehyde. As a stoichiometric agent, the solid will function much like an adsorbent bed, becoming saturated with product and requiring regeneration as was described in the Mobil patent for using zeolite. After regeneration to displace the product, the fixed bed of solid acid can be reused to produce another "batch" of arene aldehyde. One may obtain a continuous process by adapting the process strategy used in multiple, fixed-bed adsorption units so that any time, some of the fixed beds are in use, while others are in the different stages of operation (purging, regenerating, *etc*). Thus, the process design and economics of multiple fixed bed adsorption units must be used in developing a large-scale, continuous process for arene formylation.

Decorated silica. Alvaro, *et al.*, describe the synthesis of an MCM-41 that had been decorated with 1,2,2-trifluoro-2-hydroxy-1-trifluoromethyl-ethane sulfonic acid Beta-sultone.[93] This compound reacts with surface silanols by a condensation reaction to produce a tethered perfluoropropane sulfonic acid that is reactive towards acylation reaction. If this tethered perfluoropropane sulfonic acid is reactive for toluene carbonylation, then it may be a viable candidate for testing further.

Reactivity and Recovery--Reprise

We commented earlier how triflic acid was found to be important in acid catalysis[94], but the precise nature of the mechanisms by, which it works and the exact species that exist are not well known[95] despite the large effort in characterizing this system. It is appropriate that we reflect again on this conversion agent with an emphasis on how one could use it or a related acid in a commercial process to produce the arene aldehyde. The chemistry to produce either benzaldehyde or tolualdehyde, for example, involves the carbonylation of either benzene or toluene with carbon monoxide. In the case of benzene with TfOH, this reaction may proceed as an aromatic electrophillic substitution:

$$H^+ + CO \Leftrightarrow HCO^+$$
$$C_6H_6 + HCO^+ \rightarrow C_6H_5CHO + H^+$$

In the case of toluene, experimental data show conversions of nearly 100%. Presumably, the complex formed between the tolualdehyde and the acid is relatively stable, so that the back reaction is unfavorable and the complex can be dissociated by water addition. When this reaction is carried out in dry toluene, a large excess of TfOH[b] is required to drive it (TfOH:$C_6H_5CH_3$} of 10:1, mol ratio). Moreover the resulting tolualdehyde and acid complex is relatively stable. A large excess of water (13 mol water to 1 mol of acid) is required to reduce the acidity to a level that liberates the aldehyde. The resulting mixture is biphasic with an upper oil layer containing much of the aldehyde, but the process is inherently inefficient due to the large excess of TfOH that is required to drive the reaction, and the large water dilution ratio that is required to recover the aldehyde in a second phase. Although the resulting mixture will yield two liquid phases, this approach is impractical since it makes the subsequent recycle of TfOH expensive and requires the removal of significant amounts of water. Modifications of this procedure are needed to make it practical.

Super acids other than triflic may also be attractive for carbonylation. Some are claimed to be over a thousand times stronger. The $HBr/AlBr_3$ system, for example, is of interest (H_o = - 17.5[96,97] *versus* H_o=-14.1 for triflic acid, where H_o is the Hammett acidity function. While bromine is highly corrosive in the presence of water, this problem may be much less severe in the case of anhydrous systems (*e.g.*, based on dry ionic liquids). Other superacids of interest include CF_3SO_3H/SbF_5 (H_o = -16 to -18.5)[96] and HF/SbF_5 (H_o < -20)[96]. Both of these latter two systems have corrosion issues when mixed with water, but may be practical if restricted to anhydrous ionic mixtures. However, the high vapor pressure of HF coupled with its toxicity makes in unattractive as a commercial conversion agent.

What is needed are more efficient methods to carry out the reaction and subsequent recovery operations. Efficiency can be increased by finding reaction conditions that: (1) enable aldehyde production using only small amounts of TfOH or other superacids, (2) permit aldehyde liberation from the superacid without the addition of large amounts of water, and (3) facilitate the recycle of reagents and the clean recovery of product from the crude reaction mixture. The ideal system chemistry would permit both reaction and recovery from a single vessel or two, at most. Suggestions are made to improve recovery from an aqueous phase and to investigate kinetics and recovery from anhydrous mixtures of room temperature, ionic liquids (RTIL's).

Recovery from Organic/Aqueous Biphasic Mixtures. Liquid/liquid systems are affected by a wide range of factors[98,99]. We describe here a process for recovering triflic acid (TfOH) from a product mixture, but one could also imagine that the same process could be used for any member of the perfluoroalkane sulfonic acid family. Variables of interest include the mixture temperature, the presence of salting agents in an aqueous phase, and the presence of diluents in an organic phase. The patent literature[68-70] shows the use of trifluoroacetic acid as an agent to break the aldehyde/perfluoroalkane sulfonic acid complex and one may speculate that its use was to change the solubility properties of the aldehyde in the acid solution and perhaps functions like a brine, mentioned here. In the present instance, aqueous salting and the use of diluents are of particular interest because the formation of two liquid phases requires that the triflic acid-

[b] This large excess of TfOH may be required to stabilize HCO^+ in toluene solution. If so, then the use of a more polar solvent should reduce the TfOH requirements.

aldehyde adduct be broken. This means that the Gibbs energy of TfOH in the aqueous phase must be lower than that of the adduct and, similarly, the Gibbs energy of the aldehyde in the organic layer must also be lower.

Adding a non-reactive, non-polar diluent can promote the decomposition of the aldehyde-TfOH adduct and transport of the pure aldehyde into the organic layer. Organic diluents are commonly used in solvent extraction to tune solvent properties and promote more favorable conditions either to extract species, improve solvent selectivity, and/or to induce liquid/liquid phase separation[91]. The selection of an appropriate organic diluent requires some knowledge of the extraction mechanism and other phase properties (*e.g.*, interfacial tension), which may need adjustment in order to bring about the desired mass transfer and phase separation.

Similarly, the aqueous phase can be adjusted through the use of salting agents (*i. e.*, "brines") which are themselves inextractable, but which increase the ionic strength of the brine, the density differences between the two liquid phases, and their interfacial tension. In the case of TfOH, salting should also promote dipole-dipole interactions which may further induce its transfer into the brine. Because of the reactivity of TfOH and the need to recycle the acid, a salting agent such as sodium triflate, or perhaps copper(I)-triflate, are logical additives to create a brine. Aqueous salting has the further benefit of reducing the equilibrium concentrations of organic species (*e. g.*, aldehydes) in the brine, and thereby reducing aldehyde losses to it.

Assuming a suitable liquid/liquid system can be found with modest additions of water, salting agent, and diluent, then a practical TfOH recycle scheme might consist of the following steps. First, complete the reaction, then form the appropriate liquid/liquid phase. Back extract the aqueous phase to remove residual aldehyde concentrations using the organic diluent. Wash the extract with brine to remove residual TfOH from the aldehyde and thus prevent its interference with aldehyde treatment. Recover and purify the aldehyde using distillation. Recycle TfOH with distillation of water and TfOH. The TfOH will likely require a polishing step to remove residual water prior to recycle. The aldehyde may require polishing to reduce polymerization and decomposition during its distillation.

Ionic Liquids. Using anhydrous RTILs as reaction solvents, may enable higher reaction rates with the addition of less TfOH or the use of other, more effective superacids. Ionic liquids may also promote TfOH recycle, and enable product recovery without the addition of water. They have the potential to drastically simplify recycle and recovery, as well as improve reactivity and/or selectivity. Laali *et. al.*[100] discuss the use of TfOH as promoter and *n*-butylmethylimidazolium triflate ([BMIM][TfO]) room temperature ionic liquid (RTIL) as solvent. They found this mixture resulted in a mild and efficient process for the one-pot adamantylation of aromatic substrates (*e.g.*, toluene). Under optimal conditions, the reactions exhibit high *para* selectivity with little or no adamantane side-products being formed. These results suggest that a RTIL may be coupled with a suitable inert diluent to promote the decomposition of TfOH-aldehyde adducts without the addition of any water. They also demonstrated that [BMIM][TfO] could be reused. They recovered their alkylates by diethyl ether extraction.

The results of Laali *et al.*[100] suggest that toluene aldehydes may be efficiently produced using an ionic liquid without water additions. The suggested procedure is to react CO and toluene in a [BMIM][TfO] and TfOH mixture, then depressurize and recover the products, unreacted toluene, and TfOH by distillation directly from [BMIM][TfO]. It should be possible to further purify the products and reactants using fractional distillation.

The solubility of carbon dioxide in [BMIM][PF_6] is apparently higher than that of CO[101], but the data are conflicting for CO in [BMIM][TfO][102,103,104]. The solubility of CO in [BMIM][PF_6] may be an order of magnitude less than CO_2 under similar conditions. Further measurements are needed to determine CO solubilities in [BMIM][TfO]. Apparently the anion plays an important role[103].

Another modification which may be attractive is to replace catalytic TfOH with a non-volatile, catalytic triflate salt. Handy and Egrie[104] were able to catalyze aromatic nitration using ytterbium triflate as the catalyst along with N-butyl-N-methylpyrrolidinium triflate as the room temperature, ionic liquid (RTIL)

solvent and toluene as the substrate. They were also able to catalyze the same reactions using copper(II)-triflate as the catalyst. If copper triflate can be used, rather than TfOH, then only the aldehyde products, the CO, and toluene reagents will be removed by volatilization. In this simplest case, both the catalyst and ionic liquid remain in the reactor for reuse without additional handling or purification. The implementation of strategies to continuously add and remove products, CO and toluene would enable continuous operation.

Blended RTILs have not been widely studied, but it is known that RTIL mixtures often have drastically different properties from their pure constituents[105]. In some instances, the mixture properties are superior[106], but few data are available on the ability of RTIL mixtures to affect organic reactivity. It is known that reaction rates often increase when an ionic liquid is the solvent[105].

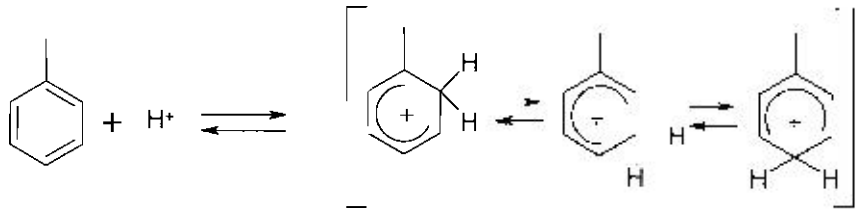
Conclusions

Arenes can be carbonylated to high conversions at room temperature under CO partial pressures of 70 atm with high regioselectivity when the catalyst is triflic acid. The reactivity and selectivity of the toluene carbonylation reaction may be adjusted by the initial ratio of acid/substrate. Carbonylation of di- and tri-substituted arenes at the same conditions in excess triflic acid showed high regioselectivities to the aldehyde addition product without rearrangement of the substrate before the carbonylation. The homologous members of the triflic acid family [$CF_3(CF_2)_nSO_3H$] were reported to show lower acid strengths than triflic acid and thus represent an opportunity to balance reactivity to form the product aldehyde against ease of separation of the acid from the aldehyde product. One patent claims to show recovery of the product aldehyde by simple vacuum distillation when a co-solvent was used. Other studies show that the addition of "salts" was effective in changing the solubility properties of the resulting liquids. Zeolites were reported to be active for arene carbonylation provided that Cu was introduced into the solid and that HCl gas was present also. Separation from the product was achieved with the introduction of a "springing" agent such as steam, CO_2, or ammonia followed by heating so as to regenerate the solid.

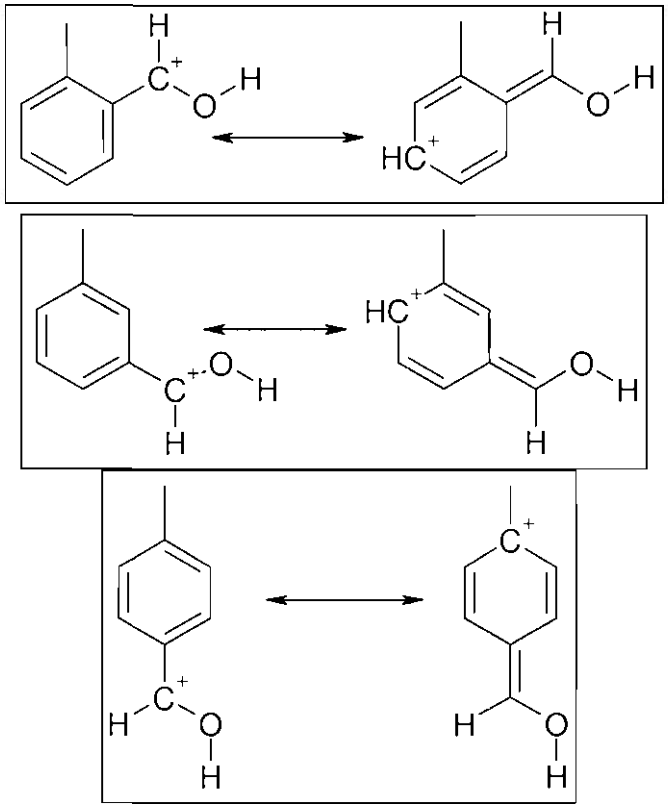
Quo Vadis

As a result of this review of the literature it appears that two paths forward are possible: 1) examine the perfluoroalkane sulfonic acid family as conversion agents for arene carbonylation (benzene or toluene) with an emphasis on the reuse of the sulfonic acid and 2) examine the reactivity and reuse of zeolites decorated with Cu and treated with HCl.

Scheme 1 Equilibria for Protonated Toluene



Scheme 2 Mesomeric forms for protonated tolualdehydes



Appendix II--Progress Reports to BP (USA) for Molecular Modeling of Acid/aldehyde Interactions

Please find in this appendix the following progress reports:

June 2005
May 2005
April 2005
March 2005
February 2005
January 2005
November 2004

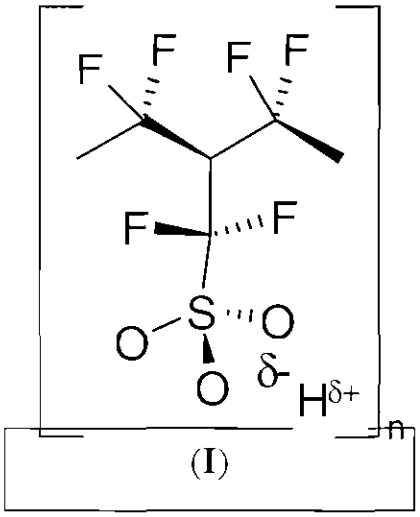
Monthly Progress Report to BP (June, 2005)
Modeling Strong Acids as Hydrocarbon Conversion Agents

Mark G. White, President
Advanced Materials WI Technologies, Inc.
1548 Lake Koinonia Drive; Woodstock, GA 30189

Executive Summary

The modeling by *ab initio* methods was continued on models of the solid acid, Nafion, to determine the tendency of these acidic species for forming adducts with *p*-tolualdehyde and Lewis bases such as dimethyl ether and diphenyl ether. These calculations showed that both ethers appear to warrant further study as extracting agents especially for displacing model compounds of coke residues on the acid sites in Nafion. Dimethyl ether appears to be a promising candidate for displacing tolualdehyde from perfluoropentane sulfonic acid.

Introduction



(I)

At the highest level of modeling, we choose Nafion, to model the interactions between the acidic sites and the Lewis bases: alkene, aldehyde, and diphenyl ether. This appears to be a very promising model system since the patent literature[c] suggests that the solid acid can be regenerated using a simple displacing agent, diphenyl ether. Moreover, the structure of Nafion is known (**I**) and the acidic site is very similar to that found in triflic acid (CF_3SO_3H). Finally, we have demonstrated that diphenyl ether is unreactive with itself in triflic acid; whereas, other ethers, such as diethyl ether and MTBE, will react in triflic acid.[d] Thus, we expect that diphenyl ether will be a good choice for a displacing agent. The identity of the "coke" which covers the surface is unknown; however, we may be able to simulate the properties of this coke from a consideration of the most probably coke formation reactions. One possible reaction is that required for the formation of the multiple alkylation products and another possible coke-forming reaction is the oligomerization of dodecene. Tasset[l] produced a coked Nafion catalyst by passing a feed rich in the olefin (tetrapropylene/diphenyl oxide = 10/1 mol/mol) so as to accelerate the coke-forming reactions.

Results

Modeling Nafion. We can imagine a simple model that captures the essential chemical properties of the solid Nafion: $CF_3(CF_2)_nCF_2SO_3H$ where n can be adjusted. As a first trial, we will thermochemical properties for the Nafion model with n = 4. The free energy of displacement of the Lewis base from the Nafion will be the criterion for judging how the value of n affects the thermochemical properties. Geometry optimizations will be completed using Hartree Fock method with a basis set at the level of 3-21(G*) and then finishing with a single point energy calculation using AM-1 in the Spartan '02 quantum mechanics package.

Lewis bases. The patent literature suggests that diphenyl oxide will remove a coke residue from the surface of Nafion. As models for the coke residue we will use the following Lewis bases: 1) 1-dodecene; 2) 1,3-dodecadiene; and 3) hexylbenzene.

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[c] Tasset, E. L., "Regeneration of Fluorocarbonsulfonic Acid Catalysts", US Patent 4,800,286 (1989).
[d] Sherman, S. C., private communication, (2002).

Extracting agents. The ethers dimethyl ether (DME) and diphenyl ether (DPE) will be considered as potential extracting agents, E, for removing models for the coke residue from Nafion.

Displacement reactions. The test reactions for these Lewis bases, L, for these extracting agents, E, will be as follows:

$$1) \quad E + Nafion/L \rightarrow L + Nafion/E; \Delta G_{rxn}^{o} = ?$$

The extracting agent, E, will be considered encouraging if the displacement reaction shows a free energy that is negative. Considering the uncertainty in the thermochemical properties of +/- 5 kcal/mole, we are expecting free energies more negative than 10 kcal/mol as very encouraging.

The results for these simulations are shown in Table 1 for the model of Nafion where n = 4. Reactions 1-4 show positive free energies of displacement (~ 10 to 12 kcal/mol) for the removal of the model coke compounds dodecene and dodecadiene using dimethyl ether (DME) and diphenyl ether (DPE) as the displacing agents. The error in these calculations is +/- 5 kcal/mol so it appears that the displacement of these types of model coke compounds is are not favored. These calculations do not consider the free energy of solution for the Lewis bases into the ethers. On the other hand, if the model for the coke is hexylbenzene, the free energies of displacement are negative (-6 kcal/mol) when using either diphenyl ether (reaction 5) or dimethyl ether (reaction 6). These results predicted by the *ab initio* method are different from those we reported last month[e] using AM-1 as the method for finding the optimum geometry where all reactions (1-6) were predicted to have negative free energies of reaction (- 6 kcal/mol for DPE and –11 kcal/mol using DME as the extracting agent).

Table 1 Thermochemical Properties for Displacement Reactions, kcal/mol

Displacement Reactions	ΔH_f^o	ΔS_f^o	ΔG_f^o
1) DPE + Nafion/dodecene --> dodecene + Nafion/DPE	8.36	-0.0065	10.29
2) DME + Nafion/dodecene --> dodecene + Nafion/DME	12.91	0.0090	10.22
3) DPE + Nafion/dodecadiene --> dodecadiene + Nafion/DPE	9.01	-0.0102	12.04
4) DME + Nafion/dodecadiene --> dodecadiene + Nafion/DME	13.56	0.0053	11.98
5) DPE + Nafion/hexylbenzene --> hexylbenzene + Nafion/DPE	-1.85	0.0155	-6.47
6) DME + Nafion/hexylbenzene --> hexylbenzene + Nafion/DME	2.70	0.0310	-6.53

It is appropriate to examine these results in more detail where we compare the free energy of formation for the small molecules (Table 2) when using either AM-1 or the *ab initio* method to arrive at the geometry optimization. Recall, that a single point energy was determined using AM-1 for each calculation after the geometry had been established using either the semi-empirical or *ab initio* method so the differences in the energy values in Table 2 reflect the differences in the optimized geometry.

Table 2 Comparison of Free Energies of Formation (kcal/mol) for Small Molecules

	AM-1	*ab initio*	literature values[f]
diphenyl ether	-3.14	6.41	34.47
tolualdehyde	-42.9	-37.18	
dimethyl ether	-67.04	-67.04	-26.99
hexyl benzene	-51.93	-51.93	-13.15
dodecadiene	-65.38	-42.84	
dodecene	-93.11	-66.96	32.96

[e] White, Mark G., Monthly Progress Report, May, 2005.

[f] Stull, D. R., E. J. Westrum, Jr., G. C. Sinke, "The Chemical Thermodynamics of Organic Compounds", J. Wiley and Sons, Inc., (1969).

The two methods predict the same free energy of formation for dimethyl ether and hexyl benzene, but significant differences are apparent in these values for predicting the free energies of diphenyl ether (~10 kcal/mol); dodecadiene (23 kcal/mol); and dodecene (26 kcal/mol), whereas the difference between the two methods in predicting the energy for tolualdehyde (5 kcal/mol) is considered insignificant. The predicted free energies of formation for perfluoropentane sulfonic acid differ by 50 kcal/mol. The literature values for the free energy of formation for hexylbenzene was −13.15 kcal/mol which is more positive than the values predicted by either method. Likewise, the literature values for the free energies of formation were more positive than the corresponding values for both ethers: dimethyl ether (~ 40 kcal/mol for DME and ~ 30 kcal/mol for DPE). Finally, for dodecene, the literature values were ~30 to 60 kcal/mol more positive than the *ab initio* and AM-1 predictions, respectively. In every case, the *ab initio* method predicted the free energies closer to that observed in the literature, but all of the predicted values were significantly different from the literature values.

We consider further these displacement reactions (Table 3) when the Lewis base is *p*-tolualdehyde (*p*-tal) and for Nafion: $CF_3(CF_2)_nCF_2SO_3H$ where n = 4. The free energy of the replacement reaction is nearly the same (-3.8 kcal/mol) when the extracting agent is dimethyl ether or diphenyl ether. This result compares well with the result reported last month (-1 to −6 kcal/mol) when the AM-1 method was used to establish the optimized geometries of all molecules.

Table 3 Nafion Model [$CF_3(CF_2)_4CF_2SO_3H$] for the Displacement Reactions of Tolualdehyde, kcal/mol

	ΔH_f°	ΔS_f°	ΔG_f°
DPE + Nafion/p-tal --> p-tal + Nafion/DPE	-4.00	-0.0008	-3.75
DME + Nafion/p-tal --> p-tal + Nafion/DME	0.55	0.0147	-3.82

These results using perfluoropentanesulfonic acid as the model for Nafion are even more interesting when we consider the possibility of using this agent as an alternative to triflic acid for toluene carbonylation. The salient question is: Can we break the acid/p-tal complex more easily when the acid is perfluoropentanesulfonic acid rather than triflic acid. To answer this question we can determine the free energy of displacement for the triflic acid system and compare those free energies to that just calculated for the pentanesulfonic acid system. These calculations were completed at the Hartree-Fock level using 3-21(G*) as the basis set for the electronic orbitals (Table 4).

Table 4 Free Energies of Reaction for Displacement of Tolualdehyde from Triflic Acid, kcal/mol

	ΔH_f°	ΔS_f°	ΔG_f°
DPE + triflic acid/p-tal --> p-tal + triflic acid/DPE	-3.70	0.0048	-5.14
DME + triflic acid/p-tal --> p-tal + triflic acid/DME	8.41	0.0169	3.35

These results suggest that tolualdehyde cannot be displaced from triflic acid by the action of dimethyl ether while this ether could be used to displace tolualdehyde from perfluoropentane sulfonic acid. Additionally, diphenyl ether is predicted to be an effective displacing agent from either triflic acid or perfluoropentane sulfonic acid. The predicted result for the use of dimethyl ether in displacing tolualdehyde from triflic acid can be compared to an experimental result that we observed several years ago[8]: dimethyl ether did not displace tolualdehyde from triflic acid. These experimental results would seem to confirm the *ab initio* prediction. If perfluoropentane sulfonic acid is active for toluene carbonylation and if the product aldehyde can be easily displaced by an ether, then the perfluoroalkane sulfonic acid might be useful as a regenerable "catalyst".

[8] White, Mark G., private notes, BP-AMOCO contract, summer, (2002).

Monthly Progress Report to BP (May, 2005)
Modeling Strong Acids as Hydrocarbon Conversion Agents

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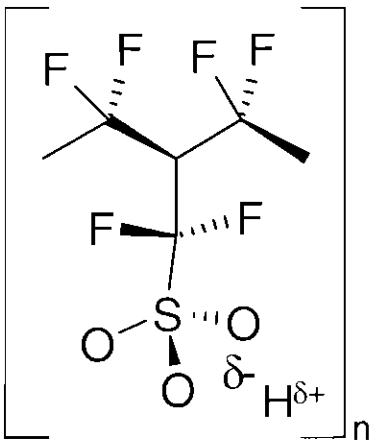
Executive Summary

The modeling by semi-empirical methods was continued on models of the solid acid, Nafion, to determine the tendency of these acidic species for forming adducts with *p*-tolualdehyde and Lewis bases such as dimethyl ether and diphenyl ether. These calculations showed that both ethers appear to warrant further study as extracting agents especially for displacing model compounds of coke residues on the acid sites in Nafion.

Introduction



(I)

At the highest level of modeling, we choose Nafion, to model the interactions between the acidic sites and the Lewis bases: alkene, aldehyde, and diphenyl ether. This appears to be a very promising model system since the patent literature[h] suggests that the solid acid can be regenerated using a simple displacing agent, diphenyl ether. Moreover, the structure of Nafion is known (**I**) and the acidic site is very similar to that found in triflic acid (-CF_3SO_3H). Finally, we have demonstrated that diphenyl ether is unreactive with itself in triflic acid; whereas, other ethers, such as diethyl ether and MTBE, will react in triflic acid.[i] Thus, we expect that diphenyl ether will be a good choice for a displacing agent. The identity of the "coke" which covers the surface is unknown; however, we may be able to simulate the properties of this coke from a consideration of the most probably coke formation reactions. One possible reaction is that required for the formation of the multiple alkylation products and another possible coke-forming reaction is the oligomerization of dodecene. Tasset[1] produced a coked Nafion catalyst by passing a feed rich in the olefin (tetrapropylene/diphenyl oxide = 10/1 mol/mol) so as to accelerate the coke-forming reactions.

Results

Modeling Nafion. We can imagine a simple model that captures the essential chemical properties of the solid Nafion: $CF_3(CF_2)_nCF_2SO_3H$ where n can be adjusted. As a first trial, we will thermochemical properties for the Nafion model with n = 4 and 9. The free energy of displacement of the Lewis base from the Nafion will be the criterion for judging how the value of n affects the thermochemical properties. Geometry optimizations will be completed using AM-1 in the Spartan '02 quantum mechanical package.

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[i] Sherman, S. C., private communication, (2002).

Lewis bases. The patent literature suggests that diphenyl oxide will remove a coke residue from the surface of Nafion. As models for the coke residue we will use the following Lewis bases: 1) 1-dodecene; 2) 1,3-dodecadiene; and 3) hexylbenzene.

Extracting agents. The ethers dimethyl ether (DME) and diphenyl ether (DPE) will be considered as potential extracting agents, E, for removing models for the coke residue from Nafion.

Displacement reactions. The test reactions for these Lewis bases, L, for these extracting agents, E, will be as follows:

$$2) \quad E + \text{Nafion/L} \rightarrow L + \text{Nafion/E}; \Delta G_{rxn}^{\circ} = ?$$

The extracting agent, E, will be considered encouraging if the displacement reaction shows a free energy that is negative. Considering the uncertainty in the thermochemical properties of +/- 5 kcal/mole, we are expecting free energies more negative than 10 kcal/mol as very encouraging.

The results for these simulations are shown in Table 1 for the model of Nafion where n = 4. All reactions show negative free energies of displacement (~ -6 and –11 kcal/mol) and it appears that dimethyl ether shows the more negative free energies for the displacement of the Lewis base by about 5 kcal/mol. These calculations do not consider the free energy of solution for the Lewis bases into the ethers, which may explain why the patent literature does not mention the use of dimethyl ether as an extracting agent.

Table 1 Thermochemical Properties for Displacement Reactions, kcal/mol

Displacement Reactions	ΔH_f°	ΔS_f°	ΔG_f°
1) DPE + Nafion/dodecene --> dodecene + Nafion/DPE	-3.20	0.0121	-6.79
2) DME + Nafion/dodecene --> dodecene + Nafion/DME	-8.32	0.0102	-11.37
3) DPE + Nafion/dodecadiene --> dodecadiene + Nafion/DPE	-2.74	0.0135	-6.76
4) DME + Nafion/dodecadiene --> dodecadiene + Nafion/DME	-7.86	0.0117	-11.34
5) DPE + Nafion/hexylbenzene --> hexylbenzene + Nafion/DPE	-2.88	0.0097	-5.77
6) DME + Nafion/hexylbenzene --> hexylbenzene + Nafion/DME	-8.00	0.0079	-10.35

We consider further these displacement reactions (Table 2) when the Lewis base is *p*-tolualdehyde (*p*-tal) and for two models of the Nafion: $CF_3(CF_2)_nCF_2SO_3H$ where n = 4 and 9. This last set of tests will determine the effect of Nafion model upon the extracting agents dimethyl ether (DME) and diphenyl ether (DPE). For both models of Nafion (n = 4 or 9), the free energy of the replacement reaction is nearly the same (-0.8 to +0.7 and –5.4 to –5.0 kcal/mol) when the extracting agent is dimethyl ether or diphenyl ether.

Table 2 Effect of Nafion Model [$CF_3(CF_2)_nCF_2SO_3H$] upon the Displacement Reactions, kcal/mol

	n	ΔH_f°	ΔS_f°	ΔG_f°	n	ΔH_f°	ΔS_f°	ΔG_f°
DPE + Nafion/*p*-tal --> *p*-tal + Nafion/DPE	=	2.33	0.0 06	-0.82	=	1.937	0.0041	0.70
DME + Nafion/*p*-tal --> *p*-tal + Nafion/DME	4	-2.79	0.0088	-5.40	9	-2.804	0.0075	-5.03

While these free energies are within the error bars of the calculation from the value of 0 kcal/mol, these results suggest that some type of ether may warrant further testing for removing the product aldehyde from the surface of a strongly acidic catalysts after reaction. Moreover, the positive values for the entropies of the displacement reaction suggest that the displacement may be favored by using temperatures above room temperature.

Monthly Progress Report to BP (April, 2005)
Modeling Strong Acids as Hydrocarbon Conversion Agents

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Executive Summary

The modeling by semi-empirical and *ab initio* methods reported last month was continued on ionic liquids derived from aluminum chloride, ethyl-methyl imidazolium chloride and HCl gas to determine the tendency of these acidic species for forming adducts with *p*-tolualdehyde and Lewis bases such as dimethyl ether, acetaldehyde, and formaldehyde. These *ab initio* calculations showed that all oxygenates modeled here appear to warrant further study as extracting agents.

Introduction

The present work is aimed at finding an extracting agent that 1) preferentially adds to the Lewis acidic $AlCl_3$, and 2) which can be removed from the Lewis acid by a simple heating protocol. The proposed extracting agents are as follows: 1) trimethyl amine, 2) dimethyl ether, 3) acetaldehyde, and 4) formaldehyde. For this report we will show the results of modeling by *ab initio* methods the extraction with dimethyl ether, acetaldehyde, and formaldehyde.

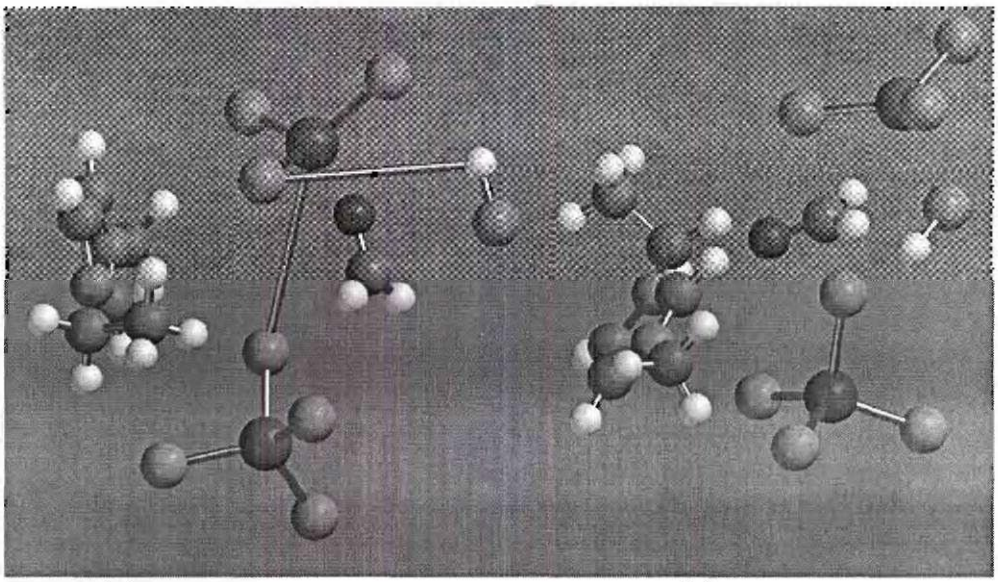
The semi-empirical calculations reported in the February 2005 report showed that free energies of extraction of –5 to –7 kcal/mol using dimethyl ether, acetaldehyde, or formaldehyde. While these results are encouraging, one must recognize the errors attending these calculations of +/- 5 kcal/mol and thus it is appropriate to reproduce these calculations using higher-level, quantum mechanical methods, such as Hartree-Fock.

Results

Ab initio model predictions-oxygenated species. Hartree-Fock calculations (Table 1) were completed to determine the usefulness of oxygenated species as extracting agents for *p*-tolualdehyde. A web site describes the accuracy of semi-empirical calculations for predicting the heats of formation.[j] All, neutral organic compounds containing C, H, N, and O atoms show an average error of 4.4 kcal/mol in the enthalpy of formation when using the PM3 method. This uncertainty rises to 9.5 kcal/mol for predicting the enthalpy of formation of organic cations (such as the imidazolium cation) when the PM3 method was used. Thus, we suggest that an extraction of the aldehyde may be possible if the predicted free energy is more negative than 10 kcal/mole for displacement *p*-tolualdehyde by the extracting agent. Apparently all reactions in table 1 appear to be possible

[j] MOPAC 2002 Manual: http:cachesoftware.com/mopac/Mopac2002manual/node650.html

according to these calculations since the smallest free energy of reaction is –22 kcal/mol. The speciation of the aluminum species appears to play no significant role as the free energies of reaction for the dimer (reaction 1) and for the anion pair (reaction 1') are predicted within the error of the calculations (*i. e.*, -22 and –25 kcal/mol are considered the same number). The reason for this result may be seen more clearly when we examine the equilibrium structures of the species involved in reactions (1) and (1'), Figure 1.



(a) (b)

The left-hand panel, 1-(a) shows the equilibrium geometry after the HF calculation and notice how long the bonds are between the Al and Cl species are which suggests that the bond order is very low in these bonds. The right-hand panel, 1-(b), shows the optimized geometry after the HF calculation beginning with the anion pair. Again, it is obvious that the length between the Al and Cl species precludes a bond. Thus, the presence of the aldehyde apparently activates the Al-Cl bond in the dimer to give the anion pair with the aldehyde inserted between the Al and Cl species of the anion pair.

Summary

From these *ab initio* calculations, we conclude that formaldehyde (HCHO), acetaldehyde (MeCHO) and dimethyl ether (MeOMe) may warrant further consideration as extracting agents for *p*-tolualdehyde/ionic liquid.

Table 1 *Ab Initio* Calculations for Oxygenates as Extracting Agents

Reactions	ΔH°_{rxn}	ΔS°_{rxn}	ΔG°_{rxn} kcal/mol
(1) HCHO + $EMIM^+Al2Cl_7^-$HCl/ptal --> ptal + $EMIM^+AlCl_4^-AlCl_3$-HCl-HCHO	13.4	0.021	-22.3
(1') HCHO + $EMIM^+AlCl_4^-AlCl_3$-HCl/ptal --> ptal + $EMIM^+AlCl_4^-AlCl_3$-HCl-HCHO	5.2	0.002	-25.0
(2) MeCHO + $EMIM^+AlCl_4^-AlCl_3$-HCl/ptal --> ptal + $EMIM^+AlCl_4^-AlCl_3$-HCl-MeCHO	-1.6	0.024	-38.2
(3) MeOMe + $EMIM^+AlCl_4^-AlCl_3$-HCl/ptal --> ptal + $EMIM^+AlCl_4^-AlCl_3$-HCl-MeOMe	0.3	0.024	-36.4

Monthly Progress Report to BP (March, 2005)
Modeling Strong Acids as Hydrocarbon Conversion Agents

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Executive Summary

The modeling by semi-empirical and *ab initio* methods reported last month was continued on ionic liquids derived from aluminum chloride, ethyl-methyl imidazolium chloride and HCl gas to determine the tendency of these acidic species for forming adducts with *p*-tolualdehyde and Lewis bases such as trimethyl amine, protonated trimethyl amine chloride [$(CH_3)_3NH^+Cl^-$] and tetramethyl amine chloride[$(CH_3)_4N^+Cl^-$]. These *ab initio* calculations showed that the protonated trimethyl amine chloride is not a viable candidate for experimental testing. The trimethyl amine and the tetramethyl amine chloride may warrant further testing by experiment to confirm the promising predictions shown here.

Introduction

The present work is aimed at finding an extracting agent that 1) preferentially adds to the Lewis acidic $AlCl_3$, and 2) which can be removed from the Lewis acid by a simple heating protocol. The proposed extracting agents are as follows: 1) trimethyl amine, 2) dimethyl ether, 3) acetaldehyde, and 4) formaldehyde. For this report we will show the results of modeling by *ab initio* methods the extraction with trimethyl amine, the hydrochloride of trimethyl amine and tetramethyl amine chloride. The choice of the tetramethyl amine chloride is motivated by its similarity to the imidazolium chloride. We found that imidazolium chloride was effective in displacing *p*-tolualdehyde from the acidic IL to form two phases, with the *p*-tolualdehyde forming the top phase.

Last month we showed conflicting data on the usefulness of trimethyl amine as a successful extracting agent for recovering *p*-tolualdehyde from the IL used in toluene carbonylation. The predictions from the semi-empirical AM1 model seemed to suggest that trimethyl amine could displace *p*-tolualdehyde from the IL; however, this result appeared to depend upon the way in which we modeled the aluminum species in the IL. That is to say, the free energy for the extraction of *p*-tolualdehyde with trimethyl amine was more negative when the aluminum species was modeled with the anion pair (-12.3 kcal/mol; $AlCl_3AlCl_4^-$) than when the species was modeled as the anion dimer (-1.2 kcal/mol; $Al_2Cl_7^-$). The negative, free energy of extraction by trimethyl amine was confirmed by a Hartree-Fock prediction (-2.2 kcal/mol) when the aluminum species was the dimer anion ($Al_2Cl_7^-$). It is necessary to establish the effect of the aluminum speciation using the Hartree-Fock calculation.

Last month we proposed the use of a protonated trimethyl amine as an extracting agent. Both the semi-empirical and *ab initio* methods suggested that the protonated trimethyl amine was not favored as an extracting agent when the aluminum species were modeled as the anionic dimer; however, the results of the semi-empirical method suggested protonated trimethyl amine could be a promising extracting agent when the aluminum species was modeled by the anion pair (-15.2 kcal/mol; $AlCl_3AlCl_4^-$). More work is needed to determine if a protonated trimethyl amine could be a candidate for further experimental work.

Results

Ab initio model predictions-methylamines. Two more Hartree-Fock calculations (Table 1) were completed to determine the effect of protonated trimethyl amine *versus* unprotonated trimethyl amine (equations 1 and 2). The *ab initio* prediction for reaction 2 was repeated starting from a different initial geometry (2') to confirm that the final structure was equilibrated and not a function of the starting geometry. The results of these three calculations (1, 2, & 2') suggested that protonated trimethyl amine is not a good extracting agent since the free energy of reaction was positive. Some comment is necessary regarding the accuracy of these calculations. A web site describes the accuracy of semi-empirical calculations for predicting the heats of formation.[k] All, neutral organic compounds containing C, H, N, and O atoms show an average error of 4.4 kcal/mol in the enthalpy of formation when using the PM3 mthod. This uncertainty rises to 9.5 kcal/mol for predicting the enthalpy of formation of organic cations (such as the imidazolium cation) when the PM3 method was used. Thus, we suggest that an extraction of the aldehyde may be possible if the predicted free energy is more negative than 10 kcal/mole for displacement *p*-tolualdehyde by the extracting agent. Thus, reaction (1) may be questionable since its free energy of reaction is only –2 kcal/mol. On the other hand, the free energies were +10-38 kcal/mol for reactions (2-4) suggesting that these reactions will not occur and thus the trimethyl amine hydrochloride is probably not a good extracting agent and it may warrant further experimental testing to confirm this prediction (*vide infra*).

A calculation was completed using the aluminum anion pair as the structure for the aluminum species (reaction 3). This prediction also showed a positive free energy of reaction (+37.7 kcal/mol) suggesting that the protonated trimethyl amine was not a candidate for extraction. We tried again to predict the efficacy of protonated trimethyl amine (reaction 4) by using the alternative structure of the aluminum anion (*i. e.*, using the anion pair: $AlCl_3HClAlCl_4^-$) and starting with $ClH^+N(CH_3)_3$ instead of HCl and $N(CH_3)_3$. The result of this additional calculation (ΔG_{rxn} = +9.6 kcal/mol) showed that the protonated trimethylamine is not a good extracting agent, regardless of the aluminum speciation and regardless of the starting amine. Thus, it appears that further work with the trimethyl amine hydrochloride may not be warranted.

We were interested in studying another amine species which was the tetra-alkyl amine chlorides as displacing agents. The motivation for this class of compounds arises

[k] MOPAC 2002 Manual: http:cachesoftware.com/mopac/Mopac2002manual/node650.html

from our experience using the dialkylimidazolium chlorides as Lewis bases and our experience with triflate salts of tetra-alkyl amines as drying agents for wet triflic acid.[1] As a final trial, we used tetramethylamine chloride as the extracting agent (reaction 5) to displace *p*-tolualdehyde and the free energy of reaction was favorable: -25.5 kcal/mol. The large negative free energy of this reaction is well outside the error anticipated for organic cations which further increases our interest in this result.

Table 1 *Ab initio* model predictions (all energies are reported in kcal/mol)

Using Hartree Fock geometry optimization/AM-1 Single Point Energy	ΔH°_{rxn}	ΔS°_{rxn}	ΔG°_{rxn}
1) $N(CH_3)_3$ + *p*-tal/EMIC/$Al_2Cl_7^-$HCl --> *p*-tal + $N(CH_3)_3$/EMIC/$Al_2Cl_7^-$HCl	-1.203	0.006178	-2.1297
2) $N(CH_3)_3$ + HCl + $EMIM^+/Al_2Cl_7HCl^-$/*p*-tal --> *p*-tal + ClH-$N(CH_3)_3$/$EMIM^+/Al_2Cl_7HCl^-$	10.748	-0.03171	15.5045
2') $N(CH_3)_3$ + HCl + $EMIM^+/Al_2Cl_7HCl^-$/*p*-tal --> *p*-tal + ClH-$N(CH_3)_3$/EMIM+/$Al_2Cl_7HCl^-$	11.013	-0.039905	16.99875
3) $N(CH_3)_3$ + HCl + $EMIM^+/AlCl_4HCl^-AlCl_3$/*p*-tal --> *p*-tal + ClH-$N(CH_3)_3$/$EMIM^+/AlCl_4HCl^-AlCl_3$	34.284	-0.022849	37.71135
4) $Cl^-H^+N(CH_3)_3$ + $EMIM^+/AlCl_4HCl^-AlCl_3$/*p*-tal --> *p*-tal + ClH-$N(CH_3)_3$/$EMIM^+/AlCl_4HCl^-AlCl_3$	11.647	0.0133	9.652
5) $Cl^-N(CH_3)^4$ + $EMIM^+AlCl_4HCl^-AlCl_3$/*p*-tal --> *p*-tal + Cl-$N(CH_3)_4$/$EMIM^+AlCl_4HCl^-AlCl_3$	-24.689	0.005217	-25.47155

Summary

From these *ab initio* calculations, we conclude that tetramethylamine chloride and trimethylamine may warrant further consideration as an extracting agent for *p*-tolualdehyde/ionic liquid, but the trimethylamine hydrochloride does not appear to hold out much promise as an extractant for this *p*-tolualdehyde/IL system.

[1] Xu, B. Q., D. S. Sood, L. T. Gelbaum, and M. G. White, *J. Catal.* **186**, 345-52 (1999).

Monthly Progress Report to BP (February, 2005)
Modeling Strong Acids as Hydrocarbon Conversion Agents

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Executive Summary

Semi-empirical and *ab initio* modeling was completed on ionic liquids derived from aluminum chloride, ethyl-methyl imidazolium chloride and HCl gas to determine the tendency of these acidic species for forming adducts with *p*-tolualdehyde and Lewis bases such as trimethyl amine, dimethyl ether, acetaldehyde, and formaldehyde. These strong acids were chosen, as they are a model for a strong Lewis and Bronsted acid pair that is often observed in commercial catalysts. The aluminum is the seat of the Lewis acidity; whereas, the HCl provides strong Bronsted acidity. Last month we showed how trimethyl amine interacted strongly with trifluoromethanesulfonic acid, a strong Bronsted acid, to show a heat of mixing that was predicted by MOPAC to be much larger than the heat of mixing predicted for triflic acid with *p*-tolualdehyde. The results shown here for the IL system also shows that trimethyl amine has a high heat of adduct formation and a negative free energy of formation, which is relevant to question of, adduct formation. When compared to the predictions for adduct formation with *p*-tolualdehyde, both the semi-empirical and the *ab initio* methods show energies of reaction that could be interpreted to suggest that trimethyl amine could displace *p*-tolualdehyde from the IL. Similar results were found for the other Lewis bases: dimethyl ether, acetaldehyde and formaldehyde. More work will be necessary to build confidence in these predictions.

Introduction

We have used chloroaluminate, ionic liquids as conversion agents for the room-temperature carbonylation of toluene.[m] These IL's were good solvents for aluminum chloride, CO, toluene and HCl so that the carbonylation could be completed at room temperature and modest CO pressures; however, the tolualdehyde product could not be easily removed from the acidic IL's. We showed that the tolualdehyde could be almost completed recovered when enough of the organic cation (disubstituted imidazolium chloride) was added to just neutralize the Lewis acid melt. Although this result was encouraging, this method of releasing the aldehyde from the IL was not commercially feasible. The results of this testing together with the most recent results suggest to us that the tolualdehyde is probably sequestered by aluminum chloride not associated with a chloride anion. The present work is aimed at finding an extracting agent that 1) preferentially adds to the Lewis acidic $AlCl_3$, and 2) which can be removed from the Lewis acid by a simple heating protocol. The proposed extracting agents are as follows: 1) trimethyl amine, 2) dimethyl ether, 3) acetaldehyde, and 4) formaldehyde. In addition,

[m] Angueira, E. J., and Mark G. White, "Arene carbonylation in acidic, chloroaluminate ionic liquids", *J. Mol. Cat. A.*, vol 227/1-2, pp 51-8 (2005)

we will examine the hydrochloride of trimethyl amine since it is functionally similar to the imidazolium chloride that was effective in displacing *p*-tolualdehyde from the acidic IL to form two phases, with the *p*-tolualdehyde forming the top phase.

Results

The acidic IL's derived from $AlCl_3$ and ethyl-methyl imidazolium chloride (2 moles $AlCl_3$/1 mole of EMIM-Cl) have been used as conversion agents for toluene carbonylation.[1,n] As with the other strong acids, water must be added to release the product from the conversion agent and in so doing, the water destroys the IL. In this work, we will calculate the free energies associated with forming an adduct between *p*-tolualdehyde and the IL and calculate the energy changes associated with adding trimethyl amine to this system either as the free amine or as the hydrochloride (Cl-H-$N(CH_3)_3$. Other candidates for the adduct include dimethyl ether, acetaldehyde, and formaldehyde.

Modeling the structure of the acidic IL's. The structure of ionic liquids is not known with certainty as it is a question still open to debate in the literature. While NMR might appear to shed light on the molecular environment, the structure and speciation of the Al anions cannot be specified yet. ^{27}Al-NMR of the neat neutral IL (stoichiometric in $AlCl_3$ and imidazolium chloride) shows a single, very sharp resonance. The IL's show two sharp resonances (one large and one very small), which are separated by 10 ppm when only a slight excess of $AlCl_3$ is present (10%). These two resonances systematically grow wider and have heights which change proportionally as excess $AlCl_3$ was added (25-50% excess) and merge into a symmetric, broad peak for IL's that are acidic (*i.e.*, 100% excess aluminum chloride). These NMR results suggest that a dynamic equilibrium exists between two structures having different Al environments. This equilibrium is speculated to be as follows:

$$[Al_2Cl_7]^{1-} \Leftrightarrow [AlCl_3\text{-}AlCl_4]^{1-}$$

While the structure of the cation is unambiguous [disubstituted ethyl-methyl-1H-imidazolium chloride ($EMIM^+Cl^-$)], two models of the anion can be imagined. One model of the acidic chloroaluminate IL's follows the line of research initially used by others in modeling the same IL that was not in contact with HCl.[o] The first model[p] is the dimeric $[Al_2Cl_7\ HCl]^{1-}$ anion and the second model is the pair[q] of aluminum species: $[HCl\text{-}AlCl_3\text{-}AlCl_4]^{1-}$. Thus, the IL will be constructed from models that we reported recently in the *AIChE Journal* and the most recent manuscript that we are now writing[4,5].

Shown in Figure 1 is the position of the *p*-tolualdehyde after the energy minimization to find the optimum geometry. Notice, here that the dimeric aluminum anion was used to construct the IL and that HCl has been placed on this dimer to produce

[n] Saleh, R., Y., "Process for making aromatic aldehyde using ionic liquids", WO-00/15594, US-6,320,083 (2000).
[o] Chandler, W. D., and K. E. Johnson, "Thermodynamic calculations for reactions involving hydrogen halide polymers, ions, and Lewis acid adducts. 3. Systems constituted from Al^{3+}, H^+, Cl^-" *Inorg. Chem.* (1999), **38**, 2050-6
[p] Angueira, E. J. and M. G. White, "Predicting the composition of acidic, chloroaluminate/EMIM-Cl ionic liquids in contact with HCl gas", accepted, *American Institute of Chemical Engineering Journal*.
[q] Angueira, E. J. and M. G. White, "Effect of reactivity of toluene carbonylation 1. Cation structure", manuscript in preparation.

the $[Al_2Cl_7HCl]^{-1}$ structure. The other model of the anion is shown in Figure 2 where the chloroaluminate anion is the pair: $[HCl\text{-}AlCl_3\text{-}AlCl_4]^{1-}$. This model is proposed to explain the ^{27}Al-NMR data that we have collected for a neat, acidic, IL where the Al/imidazolium ratio is 2 mole/mol.

Quantum mechanical methods. The systems will be modeled by the AM-1 semi-empirical model and the Hartree-Fock *ab initio* method [3-21(G*) basis set] for geometry optimization of the individual species that might be used to construct the adducts. This lower level for the basis set was chosen initially to find optimizations in a reasonable period of time. Subsequently, thermodynamic properties for these optimized geometries will be determined using AM-1 model in a single point energy calculation (SPE). The approach is to determine the geometries and energies for the systems as a sequence of stoichiometric reactions that produce the adducts of interest. We are interested in finding system(s) that will break the adduct between the IL and the product: *p*-tolualdehyde.

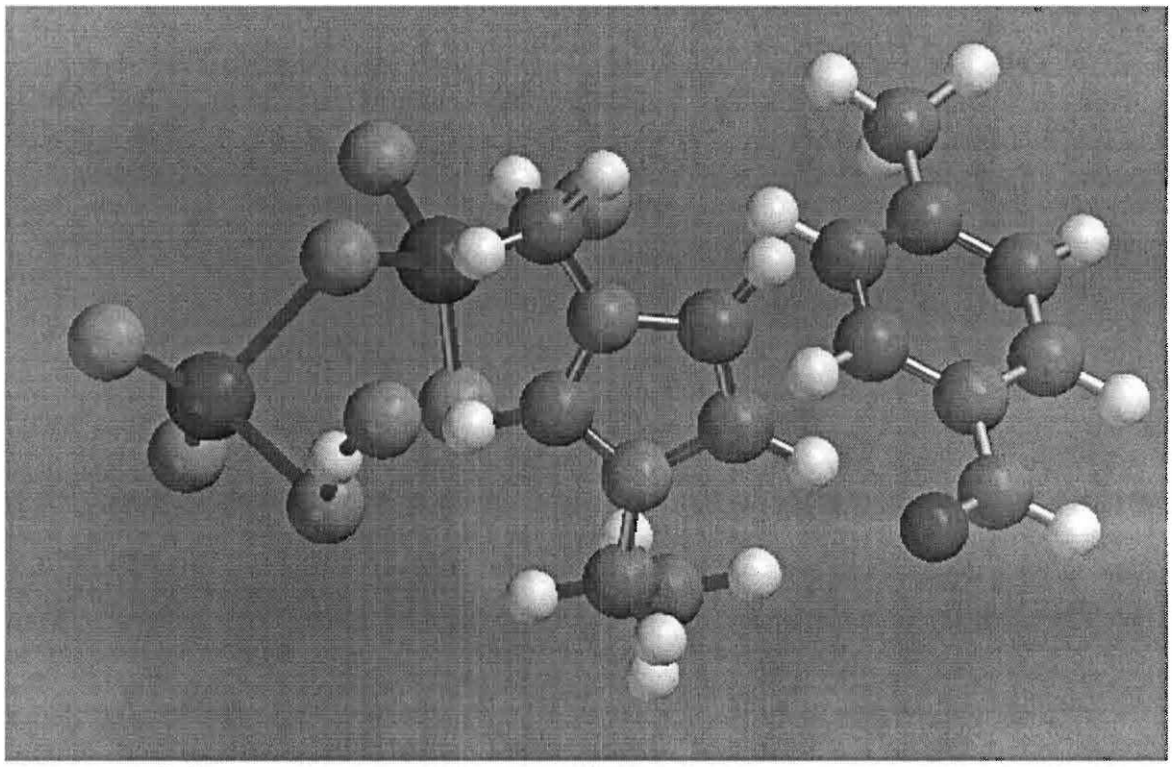

Figure 1 Model of acidic IL with *p*-tolualdehyde (dimeric aluminum anion)

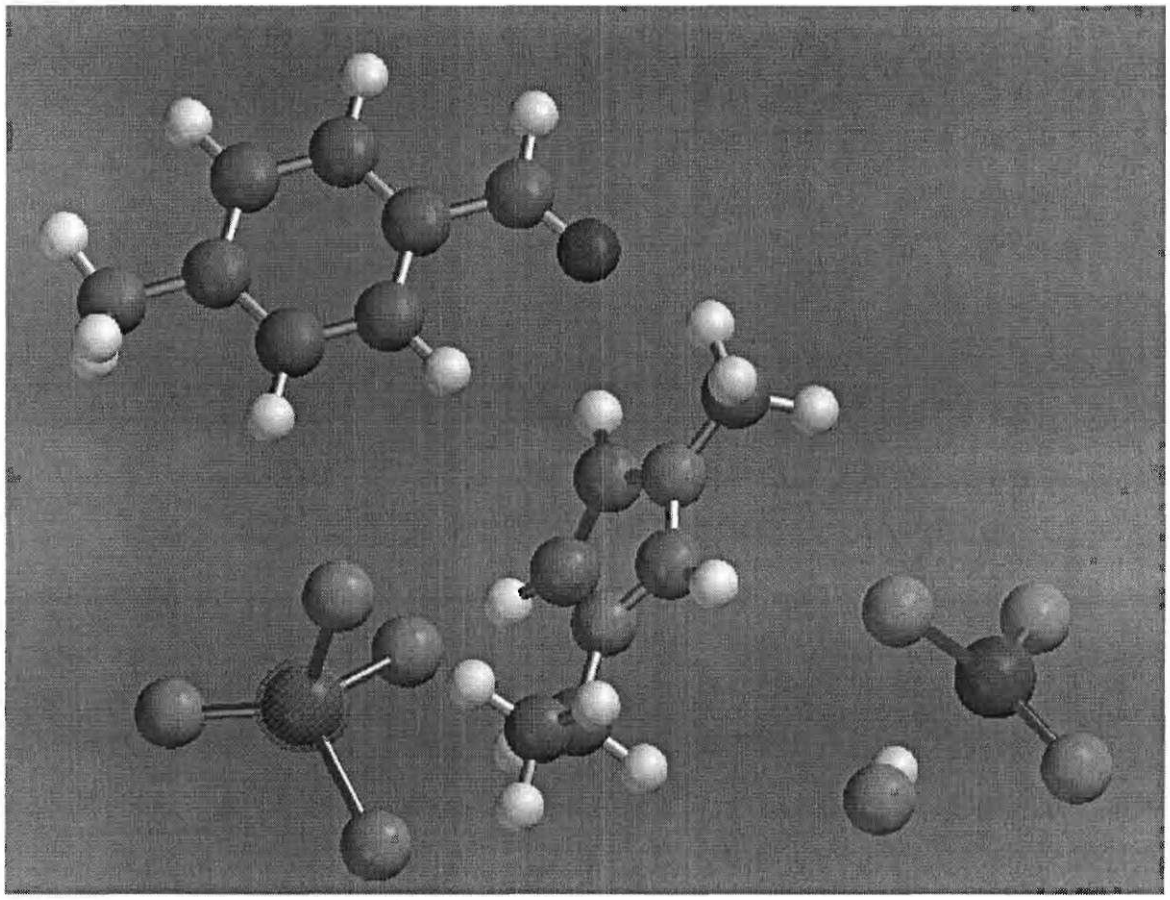

Figure 2 Model of acidic IL with p-tolualdehyde (aluminum anion pair)

We now model the interaction of the amine with the *p*-tolualdehyde/$EMIM^+$/$[Al_2Cl_8H]^-$ IL by the following reaction to determine the likelihood that the amine will displace the aldehyde from the IL. We use the stoichiometric description $[Al_2Cl_8H]^-$ to represent either the dimeric model of the aluminum anion or the aluminum pair structure.

trimethyl amine + *p*-tolualdehyde/$EMIM^+$/$[Al_2Cl_8H]^-$ →
p-tolualdehyde + trimethylamine/$EMIM^+$/$[Al_2Cl_8H]^-$ (1)

Reaction (1) is the displacement reaction that one might want to achieve so as to recover *p*-tolualdehyde from the IL. The resulting adduct with trimethyl amine could be distilled to remove the amine without causing a condensation reaction to occur which is observed when one tries to distill the *p*-tolualdehyde from a reaction mixture containing toluene.

One more consideration is the use of the hydrochloride salt of trimethyl amine as a displacing agent. The relevant reaction is:

$ClH^+N(CH_3)_3$ + *p*-tolualdehyde/$EMIM^+$/$[Al_2Cl_8H]^-$ →
p-tolualdehyde + trimethylamine-HCl/$EMIM^+$/$[Al_2Cl_8H]^-$ (2)

The use of the hydrochloride salt as a displacing agent for the tolualdehyde follows from the earlier results where the aldehyde was forced to form a second liquid phase when the imidazolium chloride was added to form the neutral IL. The formation of the adduct between the hydrochloride salt of trimethyl amine and the IL is shown in Figure 3. It was observed that the HCl molecule moved away from the amine salt as the adduct was formed with the IL.

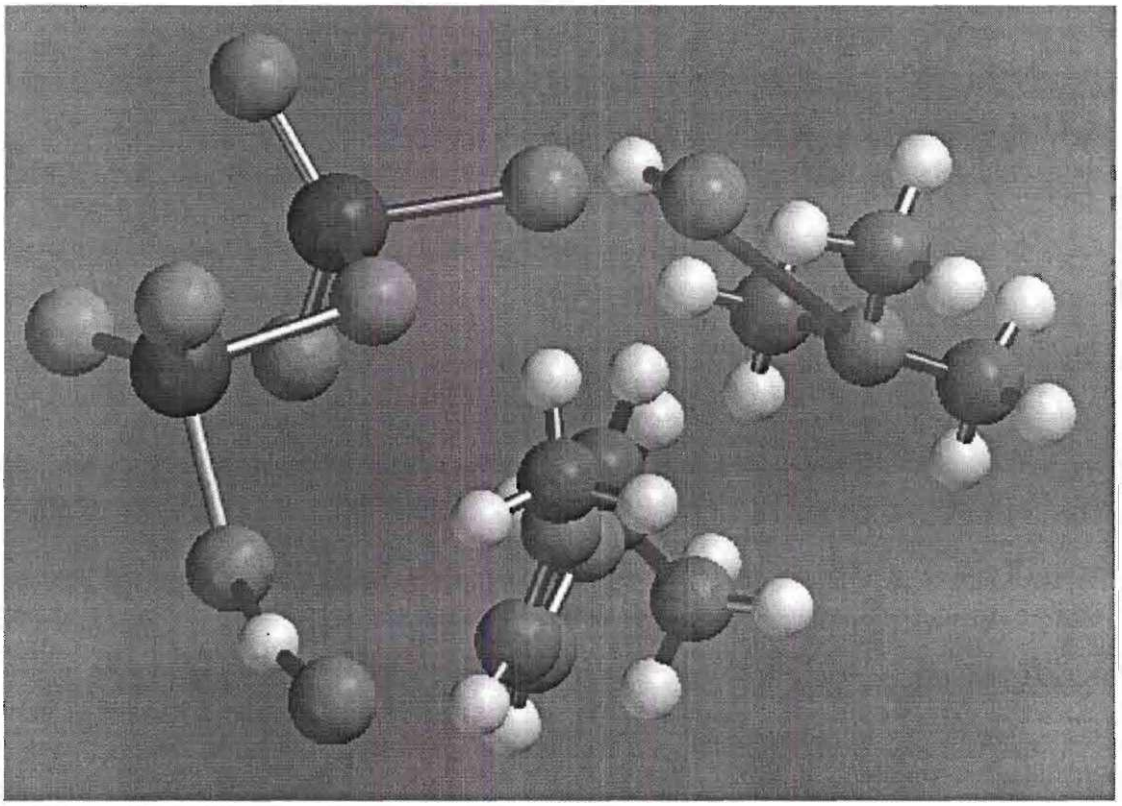

Figure 3 AM-1 model of HCl-salt of trimethyl amine with IL

We determine the energies of formation from the optimized geometries for each species in this reaction (*p*-tolualdehyde, $EMIM^+/Al_2Cl_7HCl^-$ and *p*-tolualdehyde/ $EMIM^+/Al_2Cl_7HCl^-$) and then calculate the energies of reaction and the equilibrium constants by the following:

$$\Delta H_{rxn}^{o} = \Sigma \nu_i \Delta H_f^{o}{}_i;\ \Delta S_{rxn}^{o} = \Sigma \nu_i \Delta S_f^{o}{}_i;\ \Delta G_{rxn}^{o} = \Delta H_{rxn}^{o} - T\Delta S_{rxn}^{o}$$

$$K_{equil} = \exp(-\Delta G_{rxn}^{o}/RT)$$

All of the energies were calculated in kcal/mol and the temperature used in determining the equilibrium constant was degrees Kelvin. The gas constant used in the equilibrium equation was 1.987 cal/mol-K.

Model predictions. The results of the calculations are shown in Table 1 using the semi-empirical AM-1 model (Table 1-a) and the Hartree-Fock method (Table 1-b). Consider first the semi-empirical calculations for the displacement of p-tolualdehyde from the IL,

modeled by the $[Al_2Cl_7HCl]^{-1}$ anion, as a result of the action by trimethyl amine (reaction 1 in Table 1-a). This calculation suggests that the displacement reaction is only slightly favored with a standard free energy of reaction of –1.21 kcal/mol and with a heat of reaction ~ -1.1 kcal/mol. We repeated this simulation of the displacement reaction using the pair model for the aluminum anion $[HCl\text{-}AlCl_3\text{-}AlCl_4]^{-1}$, reaction 2, Table 1-a. The pair model for the aluminum anion showed a free energy of reaction –12.3 kcal/mol with a heat of reaction of –14.7 kcal/mol. It is rather surprising that this subtle change in structure would result in a rather larger change in the thermochemical energies. Next, we repeated these calculations for displacement of *p*-tolualdehyde using the hydrochloride of trimethyl amine *via* the dimeric aluminum anion (reaction 3, Table 1-a) and the aluminum anion pair (reaction 4, Table 1-a). The hydrochloride/trimethyl amine could be a better displacing agent (ΔG^{o}_{rxn} = -20.9 *vs.* –14.7 kcal/mol) or a worse agent (+3.03 *vs.* –1.21 kcal/mol) than trimethyl amine depending upon the structure used to model the aluminum anion (dimeric or pair). These results suggest that the trimethylamine or its hydrochloride might replace the *p*-tolualdehyde as shown by the results of reactions (1-4). In the same table, we show the equilibrium constants that attend these reactions at temperatures of 150, 200, and 300 K. We purposefully picked rather low temperatures to determine if lowering the temperature would improve the replacement reaction since it is speculated that all of these reactions are exothermic. Indeed, lowering the temperature favors all of the reactions as might be expected from a consideration of the exothermic nature of all reactions; however, the benefits for lowering the temperature below room conditions are not really compelling when one considers the cost of producing these very low temperatures.

The results are also shown for using the *ab initio* method to find the optimized geometry followed by estimating the energies using a single point energy calculation by the AM-1 method (Table 1-b) for reactions (1) and (3). The Hartree-Fock calculations require ~ 15 hours each to complete on a laptop computer (IBM T42). The Hartree-Fock calculations appeared to reproduce values of the thermochemical energies predicted by AM-1 geometry optimization for reaction (1) but not for reaction (3). That is, the free energy for reaction (1) was predicted to be –1.2 and –2.1 kcal/mol using either AM-1 or Hartree-Fock methods, respectively, to find the optimized geometry. On the other hand, the free energies predicted for reaction (3) by the AM-1 and Hartree-Fock methods were +3 and +16 kcal/mol. These latter results suggest that both methods confirm that the HCl-trimethylamine may not be a good extracting agent if the aluminum speciation is that of a dimer $(HCl\text{-}Al_2Cl_7)^{-1}$ rather than the pair $(AlCl_3\text{-}HCl\text{-}AlCl_4)^{-1}$. Additional calculations for reactions (2) and (4) will be completed at the Hartree-Fock level.

Additional calculations were completed on the following extracting agents: dimethyl ether (MeOMe), acetaldehyde (CH_3CHO), and formaldehyde (HCHO) using the AM-1 semi-empirical method (Table 2). The extracting reactions were as follows:

MeOMe + *p*-tolualdehyde/$BMIM^+/Al_2Cl_7HCl^-$ →
p-tolualdehyde + MeOMe/$BMIM^+/Al_2Cl_7HCl^-$ (5)

MeCHO + *p*-tolualdehyde/$BMIM^+/Al_2Cl_7HCl^-$ →

$$p\text{-tolualdehyde} + MeCHO/BMIM^{+}/Al_2Cl_7HCl^{-} \quad (6)$$

$$HCHO + p\text{-tolualdehyde}/BMIM^{+}/Al_2Cl_7HCl^{-} \rightarrow p\text{-tolualdehyde} + HCHO/BMIM^{+}/Al_2Cl_7HCl^{-} \quad (7)$$

These results show negative free energies for the reactions (-5 to –7 kcal/mol) described by equations 5-7 thereby suggesting that these three species might be additional extracting agents to consider for further calculations using the Hartree-Fock method.

Summary

Taken together, these studies suggest encouragement for more work using higher-level calculations and possibly some experimental measurements at a later date to confirm the most favorable results.

Table 1-a AM-1 calculations for trimethylamine and *p*-tolualdehyde with $EMIM^+/[Al_2Cl_8H]^-$

		kcal/mol	cal/mol-K	kcal/mol	K @ T, K		
Using AM-1 geometry optimization		ΔH°_{rxn}	ΔS°_{rxn}	ΔG°_{rxn}	150	200	300
1	$N(CH_3)_3$ + *p*-tal/$EMIM^+/Al_2Cl_7^-HCl$ --> *p*-tal + $N(CH_3)_3/EMIM^+/Al_2Cl_7^-HCl$	-1.067	0.000929	-1.20635	5.73E+01	2.34E+01	9.56E+00
2	$N(CH_3)_3$ + p-tal/$EMIM^+/AlCl_4AlCl_3^-HCl$ --> *p*-tal + $N(CH_3)_3/EMIM^+/AlCl4^-HClAlCl_3$	-14.688	-0.015699	-12.33315	9.35E+17	4.17E+12	1.86E+07
3	$N(CH_3)_3$ + HCl + $EMIM^+/Al_2Cl_7HCl^-$/ptal --> *p*-tal + ClH-$N(CH_3)_3/EMIM^+/Al_2Cl_7HCl^-$	-0.196	-0.021498	3.0287	3.86E-05	3.28E-05	2.78E-05
4	$N(CH_3)_3$ + HCl + $EMIM^+/AlCl4HCl^-AlCl_3$/ptal --> *p*-tal + ClH-$N(CH_3)_3/EMIM^+/AlCl4HCl^-AlCl_3$	-20.896	-0.038147	-15.17395	1.29E+22	3.15E+14	7.70E+06

Table 1-b Hartree-Fock calculations for trimethylamine and *p*-tolualdehyde with $EMIM^+/[Al_2Cl_8H]^-$

		kcal/mol	cal/mol-K	kcal/mol	K @ T, K		
Using HF geometry optimization/AM-1 SPE		ΔH°_{rxn}	ΔS°_{rxn}	ΔG°_{rxn}	150	200	300
1	$N(CH_3)_3$ + *p*-tal/EMIC/$Al_2Cl_7^-HCl$ --> *p*-tal + $N(CH_3)_3/EMIC/Al_2Cl_7^-HCl$	-1.203	0.006178	-2.1297	1.27E+03	4.62E+02	1.69E+02
3	$N(CH_3)_3$ + HCl + $EMIM^+/Al_2Cl_7HCl^-$/*p*-tal --> *p*-tal + ClH-$(CH_3)_3/EMIM^+/Al_2Cl_7HCl^-$	12.01	-0.0303	16.55	7.55E-25	1.79E-20	4.26E-16

Table 2 Free Energies of Reaction

Extracting agent	ΔG°_{rxn}, kcal/mol
MeOMe	-5.25779
MeCHO	-6.99244
HCHO	-6.23654

Monthly Progress Report to BP (January 2005)
Modeling Strong Acids as Hydrocarbon Conversion Agents

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Executive Summary

High level, *ab initio* modeling was completed on small ensembles of HF/BF_3 in an attempt to predict with higher accuracy the free energies of mixing. Some improvements in accuracy were realized but not commensurate with the much longer time required for the calculations. It is simply not feasible to model liquid mixtures of product aldehyde with HF/BF_3 using *ab initio* methods. The AM-1 semi-empirical method was used to predict the free energy and enthalpy of mixing for ternary mixtures of HF/BF_3/p-tolualdeyde as to study the variation in these energy functions as HF was removed from the mixture. This modeling was motivated by the commercial practice of first removing HF from the product mixture before the aldehyde was distilled from the remaining liquid. Our modeling showed that non-monotonic changes in the free energy and enthalpy of mixing attended the stepwise removal of HF from the mixture. For certain changes in the concentrations of HF in the mixture, the free energy actually increased suggesting that phase separation could attend these variations in the HF concentrations. Finally, we show some previous MOPAC estimates of the heat of mixing that were completed so as to 1) understand calorimetric data we had compiled for the mixing of triflic acid with such bases as water, p-tolualdehyde and tetrahydrofuran, and 2) to motivate additional consideration for replacing water as a work-up agent with other agents such as trimethyl amine, acetic acid, and carbon dioxide.

Results

Ab initio modeling. Our earlier attempts were less than successful in modeling the thermodynamic properties of HF/BF_3 mixtures using the semi-empirical molecular modeling programs such as AM-1, PM3, and MNDO. The enthalpies of free energies of formation for HF, BF_3 and HF/BF_3 mixtures were estimated using the Hartree-Fock, *ab initio* method (Table 1). The basis set was the Slater type orbitals (STO-3G). These calculations were completed for ensembles of 1, 2, 3, and 4 molecules of HF, BF_3 and the pair HF/BF_3. The predicted enthalpies of these ensembles increased slightly with increasing size of the ensemble (n = 1→

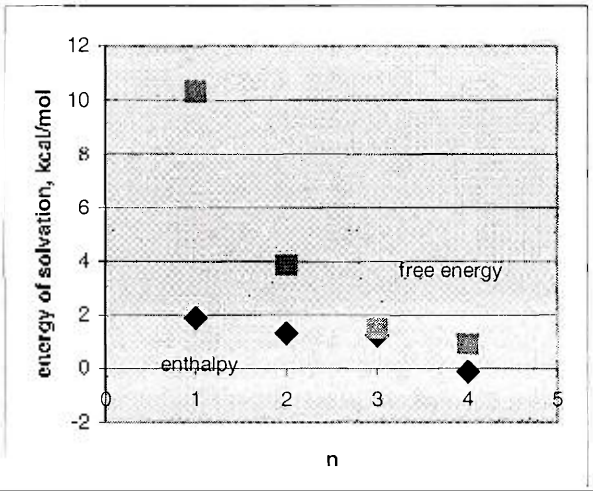

Figure 1 Energies of Solvation

4); whereas the predicted free energies of formation increased with increasing ensemble size (Table 1). However, the energies of solvation of HF with BF_3 decreased with increasing ensemble size (Table 1 and Figure 1). As the ensemble size increased to 3 and 4 molecule pairs, the enthalpy of solvation became slightly negative and the free energy of solvation became a value near 1 kcal/mol.

We may compare these estimates of liquid properties with that reported in the literature. The literature value[r] for the free energy of liquid BF_3 at 298 K is –57.5 kcal/mol; whereas, the estimate value of this free energy 10-12 kcal/mol. Compare this estimate with that we reported from the semi-empirical methods where the free energy of liquid BF_3 was estimated to be –285 kcal/mol.[s] The Hartree Fock, *ab initio* method appears to come closer to the literature value for the free energy of liquid BF_3 than that predicted by the semi-empirical method; however, the *ab initio* value is still not very good. Moreover, the CPU time required is 1 hour to model these ensembles having 4-8 molecules. The time required for modeling becomes very long when tolualdehyde is added to this mixture.

Table 1 Predicted Thermodynamic Properties of HF, BF_3, and HF/BF_3

Enthalpy, kcal/mol: n =	1	2	3	4
BF_3	10.66	11.21	11.39	11.82
HF	7.287	7.772	7.919	8.974
HF-BF_3	19.85	20.30	20.86	20.69
$\Sigma\nu_i\Delta H_i/n$	1.903	1.319	1.253	-0.112
Free Energy, kcal/mol	1	2	3	4
BF_3	-7.442	-2.260	1.136	2.131
HF	-5.09	-0.470	1.795	2.548
HF-BF_3	-2.20	5.00	7.546	8.372
$\Sigma\nu_i\Delta G_i$	10.34	3.87	1.52	0.923

AM-1 Modeling. We attempted to model the toluene carbonylation process produced from the action of HF/BF_3 by first predicting thermodynamic properties of mixtures of HF/BF_3 with *p*-tolualdehyde at room temperature and pressure. One of the goals of this modeling was to understand how the enthalpy and free energy of the mixture changed with the HF content at constant values of the BF_3 content. To accomplish this goal, we choose an ensemble that contained 1 molecule of *p*-tolualdehyde, 4 molecules of BF_3 and n molecules of HF for which n varied between 0 and 5. This ensemble size was motivated from a description of the process by which toluene was carbonylated by HF/BF_3 (*vide infra*).

[r] Lange's Handbook of Chemistry, 12TH Edition, Ed. by J. A. Dean. McGraw-Hill Book Company, (1979).
[s] White, M. G., Monthly Progress Report to BP-Amoco (November 2004), "Modeling Strong Acids as Hydrocarbon Conversion Agents".

Consider the description of the HF/BF_3 process that is the conversion agent for carbonylation of toluene to form tolualdehyde.[t] The product aldehyde is a weak Lewis base owing to the lone-pairs of electrons on the oxygen. All of the toluene substrate is combined with all of the HF and only part of the BF_3 required to initiate the reaction. To this mixture is added CO gas so as to convert a part of the toluene to tolualdehyde by a stoichiometric reaction with BF_3. This product mixture is then combined with the remaining BF_3 to complete the reaction. The HF/substrate ratio is 5-6 mol/mol toluene; whereas, the BF_3/substrate ratio is 1-1.5 mol/mol toluene and the reaction is kept at 273 K. A catalytic cycle for this system can be produced provided that the HF and BF_3 is separated from the product mixture which may be achieved in a distillation apparatus where the temperature of the mixture is raised to 110°C. At this temperature, the vapor pressure of the system is about 3 atm. The complex between the aldehyde product and the HF/BF_3 is broken as a result of the high vapor pressure of the acid components. Usually, HF is flashed from the mixture at relatively low temperatures, leaving the complex between the aldehyde/BF_3 to be broken at higher temperatures. Two distillation columns can be used so as to remove HF in the first column followed by a second column to break the Lewis acid/Lewis base complex. Therefore, one could model the breaking of the product aldehyde from the HF/BF_3 conversion agent by constructing an ensemble of 2 tolualdehyde, 3 BF_3 molecules (BF_3/substrate = 1.5), 10 HF molecules.

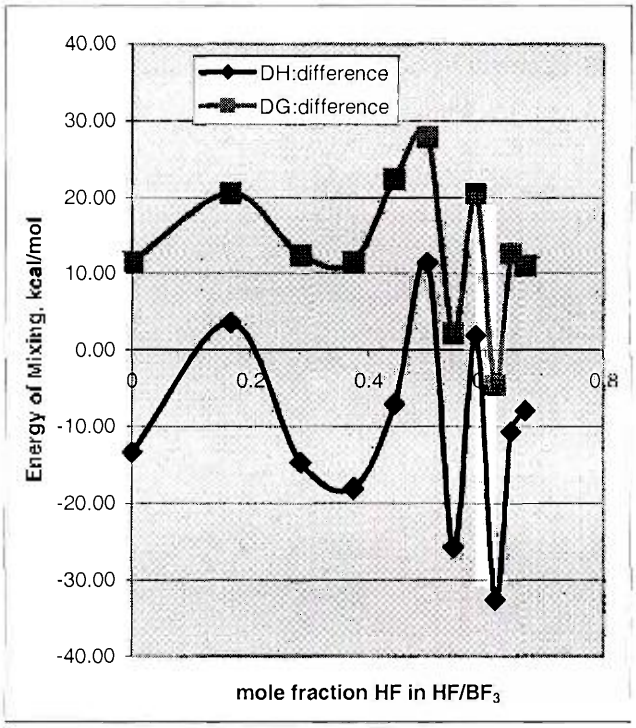

Figure 2 Enthalpy and Free Energy of Mixing

The results of this modeling efforts (Figure 2) show that the enthalpy and free energy of mixing *p*-tolualdehyde with HF/BF_3 experience maximum and minimum values for varying the HF content of the mixture on the range of HF mole fraction between 0 (all HF distilled) and 0.67 (starting composition coming from reactor where HF/BF_3 ratio = 10 moles/3 moles with 2 moles of p-tolualdehyde). Previous comparisons of predicted free energies and observed values for liquids show that the AM-1 program does not predict well the absolute value of the free energy of a liquid. However, the change in free energy as one changes the composition may be accurately predicted using this program. If we use the relative values of the free energy of mixing in relation to the values at $x_{HF} = 0$, then it may be possible to interpret these results.

[t] Fujiyama, *et al.*: "Process for producing *p*-tolualdehyde", US Patent 3,948,998 (1976); "Process for purifying *p*-tolualdehyde" US Patent 3,956,394 (1976); "Method for decomposing an aromatic aldehyde-hydrogen fluoride-boron trifluoride complex", US Patent 3,988,424 (1976);

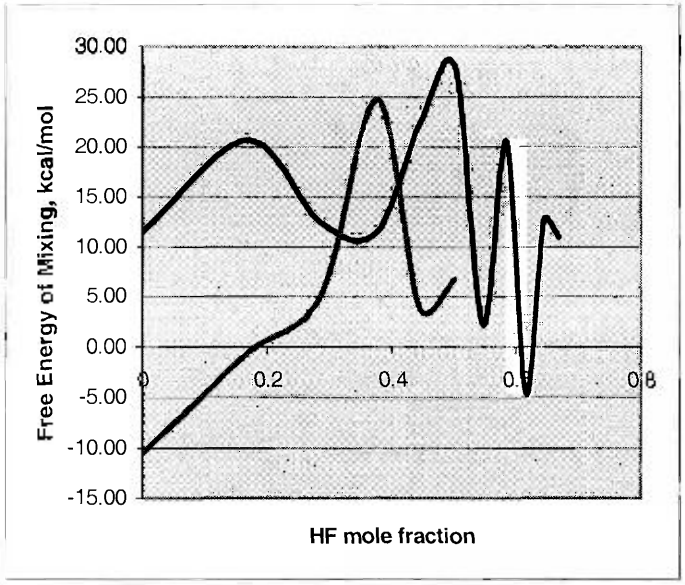

Figure 3 Relative Free Energy of Mixing

One interpretation of these results is that a phase separation may be possible as HF is distilled from the mixture, which is accomplished in the first distillation column in the continuous product recovery system. Consider the relative values of the free energy of mixing computed for this system as in Figure 3, blue line. Here, we consider only the changes in the free energy of mixing equal to zero for a mixing containing no HF. The free energy of the model fluid containing the maximum amount of HF is also very near zero (X_{HF} = 0.71). This composition would correspond to that of a fluid exiting the carbonylation reactor for which HF would be first removed by distillation. These predictions suggest that phase separation could occur as more HF was removed by distillation when the $\Delta G_{,mixing}$ becomes positive. The oscillatory nature of these predicted free energies of mixing might be attributed to the small ensemble size that was used in developing the model. We repeated this calculation using a smaller ensemble size (1 *p*-tolualdehyde, 4 BF_3 molecules and n HF molecules, n = 0 → 5). These predictions are shown in Figure 4 for the enthalpy and free energy of mixing. The wide oscillations that were present in the earlier predictions are not present here, even though the ensemble size is smaller. However, it is clear that some positive increases in the free energy of mixing can be observed as the HF is removed from mixture (positive relative change in free energy as X_{HF} decreased from 0.45 to 0.28; 4 → 3 molecules of HF with 4 molecules of BF_3).

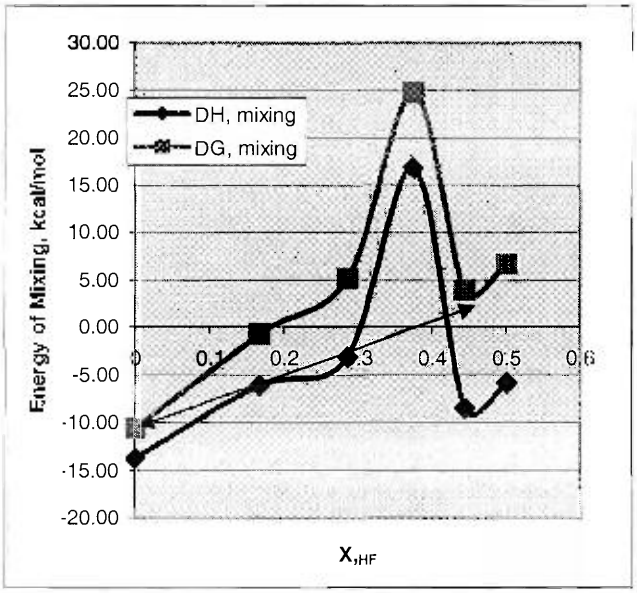

Figure 4 Free Energy and Enthalpy of Mixing

As before, we cannot literally interpret the behavior of these systems based upon the predicted free energies since the AM-1 program shows substantial errors in predicting the absolute free energy. However, if we assume that this model accurately represents the changes in the free energy with HF composition, then again a phase separation is suggested by the sudden rise in free energy with decreasing HF content. This relative maximum in the free energy curve will result in phase splitting even when the free energies in the neighborhood are all negative![u] The implication of phase splitting is that two phases may be formed, one rich in HF, the other depleted in HF. For the present case we illustrate how a mixture having an HF mole fraction on the range of 0 → 0.5 could split into two phases; one phase shows no HF and the other shows an HF composition of 0.5. So, if the mixture had an X_{HF} = 0.55 and some HF was removed in a distillation tower, these predictions suggest that two phases could result if the HF mole fraction was reduced to a value less than 0.5. *P*-tolualdehyde could be distributed between the two phases where it would be removed by distillation in the second tower. If we again take the relative free energies as true, then the more stable phase will be the one that has no HF since its free energy is –11 kcal/mol whereas the free energy of the phase where X_{HF} = 0.5 is +4 kcal/mol.

We now compare the predicted free energies of mixing for the two cases considered here (Figure 3--blue line: 2 *p*-tolualdehyde molecules; 3 BF_3 molecules; n HF molecules, n = 0→10; and red line: 1 *p*-tolualdehyde molecule; 4 BF_3 molecules; n HF molecules, n = 0→5). Both systems suggest that removing some of the HF from the reaction products could benefit the further separation of *p*-tolualdehyde from BF_3. BF_3 alone will not catalyze the further reaction of aldehyde with toluene to form di- and tri-tolymethane; therefore, heating this liquid to higher temperatures will not form condensed products.

MOPAC modeling for heats of mixing. Several years ago we modeled the heats of mixing (Figure 5) for binary systems of triflic acid (CF_3SO_3H) and agents such as 1) *p*-tolualdehyde, 2) water, 3) tetrahydrofuran (THF), 4) carbon dioxide, 5) acetic acid, and 6) trimethyl amine. For *p*-tolualdehyde, water, and THF we could compare these predictions experimental values for the heat of mixing that we determined in a very crude device. The MOPAC software over-predicted the heats of mixing for the *p*-tolualdehyde/triflic acid (5-8 kcal/mol acid) and the water/triflic system (5-10 kcal/mol acid); whereas, the model slightly under-predicted the heat of mixing for THF/triflic acid system (3 kcal/mol acid). Additional predictions showed that very high heats of mixing could be expected for mixing triflic acid with trimethyl amine and thus, one might expect that trimethyl amine could be used to displace *p*-tolualdehyde from triflic acid. Heats of mixing comparable to that for *p*-tolualdehyde/triflic acid were observed for THF/triflic acid and acetic acid/triflic acid. This last result is somewhat surprising *given that acetic acid is an acid*! More work will be necessary to convince us of this predicted result for acetic acid.

[u] Van Ness, H. C., and M. M. Abbot, Classical Thermodynamics of Nonelectrolyte Solutions, with Applications to Phase Equilibria, McGraw-Hill Book Company, 1982; page 388, Figure 6-37.

The results shown here encourage us to repeat these predictions using AM-1, PM3, and MNDO so as to obtain the free energies of mixing for all of these compounds except CO_2.

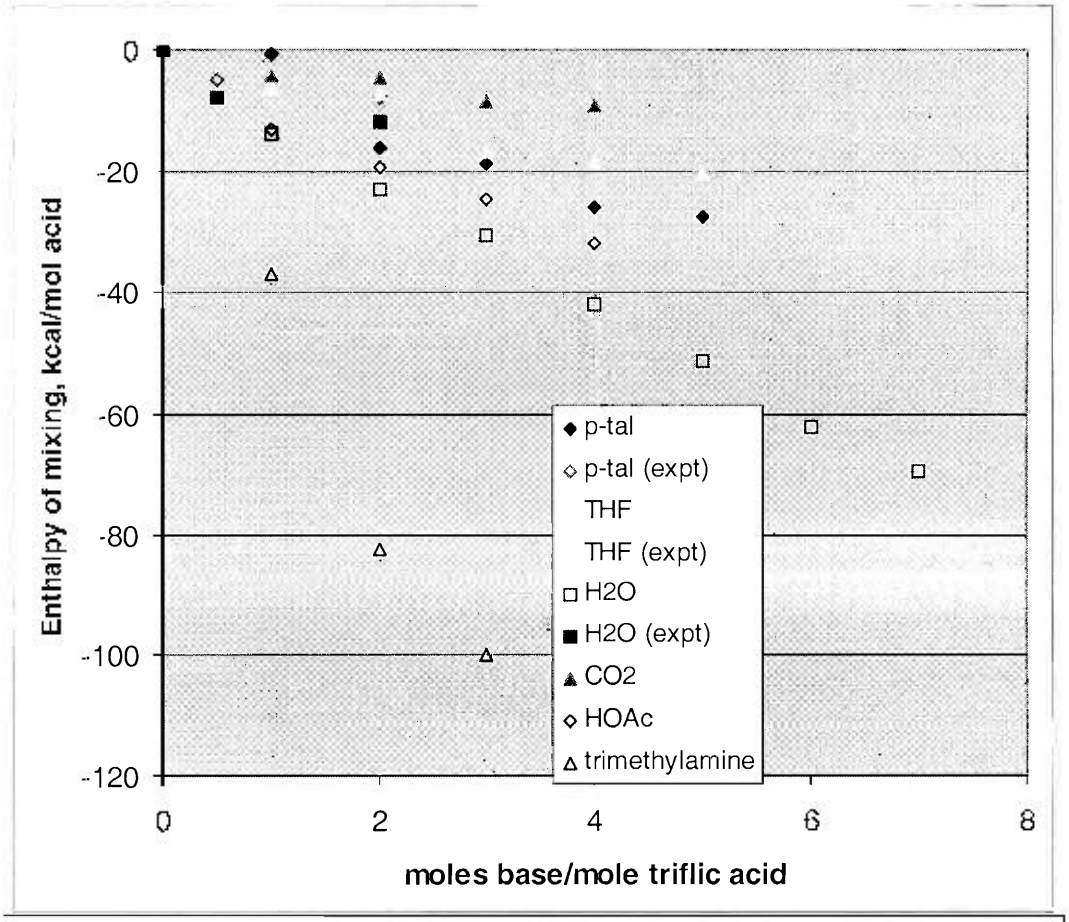

Figure 5 MOPAC predictions for binary systems of triflic acid and bases

Monthly Progress Report to BP-Amoco (November 2004)
Modeling Strong Acids as Hydrocarbon Conversion Agents

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Executive Summary

Preliminary calculations were completed to model the enthalpy and free energy of condensed phases in preparation to the modeling of strong interactions between molecules with surfaces or in liquid solutions. This project focuses on the science that governs the formation of an adduct between the product and the catalytically active ensemble. In particular, we seek to model these systems with the goal to understand how this complex can be altered either by change in reaction conditions and/or with the addition of an agent so that the catalytic ensemble can be regenerated using commercially-viable methods. Most of the target systems of this study will be condensed phases.

Introduction and Background

Predictions. Titan suite software programs (Wavefunction, Inc. and Schrödinger, Inc.) were used to predict the thermochemical properties of selected species by the following approximations: 1) AM1, a semi-empirical molecular orbital model. These models were used to perform equilibrium geometry calculations and/or single point energy (SPE) determinations. The heats of mixing for binary mixtures were estimated from a consideration of the heats of formation of the species using AM1.

Predictions—Enthalpy of mixing of binary mixtures. We attempted to model the enthalpy of mixing using AM1 to determine the enthalpy of formation for the pure components and for the mixtures that had been subjected to energy minimization to obtain the optimum geometry. Thermodynamic principles were used to express the heat of mixing in terms of measured temperature rise in the calorimeter and the heat capacities of the components of the mixture. Consider the definition of the heat of mixing for a non-ideal mixture:

$$\Delta H^{\circ}_{mixture} = \Sigma \Delta H^{\circ}_{i} + \Delta H^{\circ}_{mixing}$$

where, $\Delta H^{\circ}_{mixture}$ = enthalpy of the entire mixture,
ΔH°_{i} = enthalpy of pure component i
$\Delta H^{\circ}_{mixing}$ = heat of mixing for this mixture

D. Results – Effects of Solvation upon Enthalpy and Free Energy

The quantum mechanics programs predict the properties of the species as an isolated molecule. However, the systems that must be simulated are condensed phases where the interactions between the molecules/species are significant and contribute the total energy of the system. We may include the effects of intermolecular forces by modeling ensembles of the species so as to estimate the energies of these interactions. Consider the data in table 1 where we show the enthalpy of formation for benzaldehyde and water molecules in ensembles numbering from 1 to 12 molecules. These table entries were generated by dividing the enthalpy of formation of an ensemble containing *n* molecules reported by the quantum mechanics program by the number of molecules in the ensemble, *n*. Notice how the enthalpy of formation for the isolated molecule is less negative than that predicted for the same molecule when surrounded by like neighbors. That is, the standard state, enthalpy of formation for water as an ideal gas molecule is estimated

to be -59.24 kcal/mol; whereas, the reported enthalpy of formation for gaseous water at 298 K is -57.796 kcal/mol.[107] For larger ensembles of water, the AM1 software

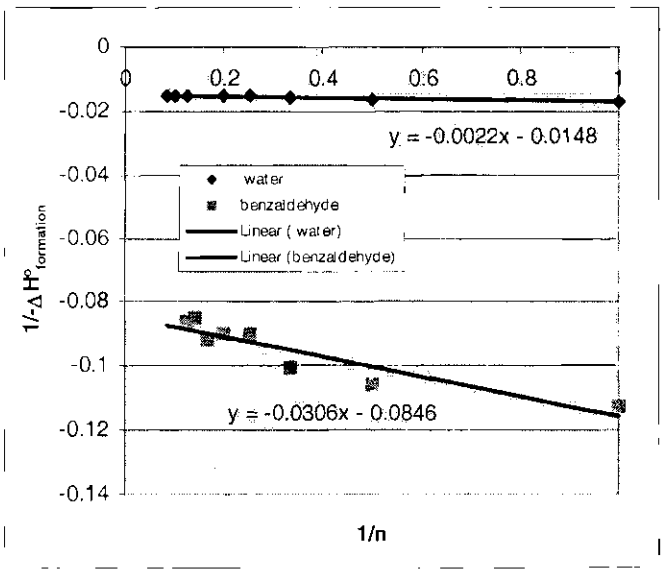
Table 1 Enthalpy of formation for ensembles of pure compounds, kcal/mol

n	1	2	3	4	5	6	7	8	10	12
water	-59.24	-61.74	-64.01	-65.22	-65.72	-61.959		-66.47	-66.88	-67.34
benzaldehyde	-8.923	-9.5095	-9.941	-11.145	-11.1058	-10.8758	-11.7836	-11.6403		

predicts values for the heat formation that are more negative, up to a value of -67.34 kcal/mol for an ensemble of 12 water molecules. The heat of formation of liquid water at 298 K is -68.32 kcal/mol.[1] The increased negative, enthalpy of formation is due to solvent effects that appear to be modeled well by AM1 *for water*. We attempted to use the same approach to model the enthalpy for liquid benzaldyde. The same trend was observed for ensembles of benzaldehyde numbering from 1 to 8 molecules where the ΔH°_{f} for the isolated molecule is –8.923 kcal/mol (literature value for gas at 298 K = -9.57 kcal/mol); whereas the ΔH°_{f} for an ensemble of 8 molecules is –11.64 kcal/mol. If we had unlimited computing power, then even better estimates of the heat of formation of liquid water and liquid benzaldehyde could be developed by considering increasing larger ensemble sizes. However this approach is not feasible owing to the time required to complete the calculation.

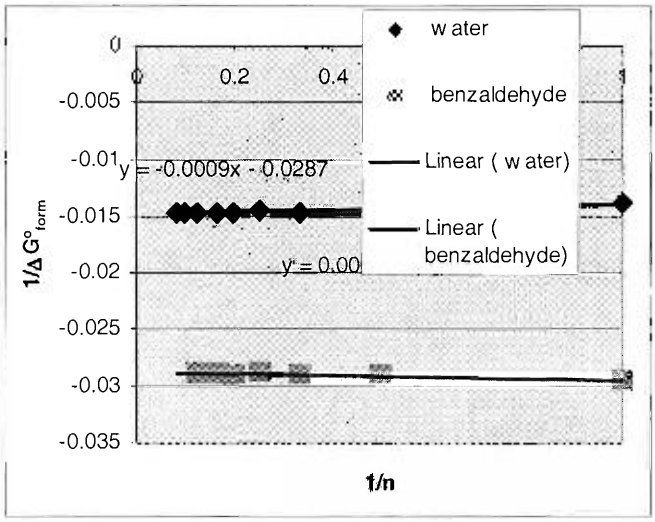
We conjecture that the properties of an ensemble of an infinite size (*i. e.*, the liquid) can be estimated from the properties of a much smaller size by a proper extrapolation. Consider, a plot of the data from Table 1 which examines the behavior of the inverse heat of mixing versus the inverse of the ensemble size, 1/n. A liquid would be modeled by $n \rightarrow \infty$ or $1/n \rightarrow 0$, so that the heat of formation for the liquid would be the value of enthalpy of the pure component at the intercept (Figure 1). From these data, we estimate that the enthalpy of pure, benzaldehyde at 298 K is –11.82 kcal/mol; whereas, liquid water is predicted to have ΔH°_{f} = -67.3 kcal/mol.

We can compare these predictions with literature data for the thermodynamic properties of pure compounds. For liquid water, the observed heat of formation is -68.32 kcal/mol; whereas, we predicted it to be -67.51 kcal/mol by extrapolating from calculated heats of formation for ensembles numbering 2, 3, 4, 5 8, and 12 molecules. If we include ensembles having only 4, 8, and 12 water molecules, the estimated ΔH_{f}° = -68.21 kcal/mol which is 0.11 kcal/mol less than the literature value. By the same method we can compare the estimated heat of formation for liquid benzaldehyde (-11.82 kcal/mol) and compared this value with the literature value (-21.23 kcal/mol). Thus, the AM-1 program produced good estimates of the enthalpy of formation for liquid water but it did not predict well the enthalpy of formation for liquid benzaldehyde.



Now consider the same approach to estimate the free energies of formation for liquid water and benzaldehyde at 298 K (Figure 2). Here we show that the extrapolated values at 298 K for the liquids are as follows: water, -67.567 kcal/mol, and for benzaldehyde, -34.84 kcal/mol. The extrapolation of the water values appears to be a smooth one whereas the data for the free energies of the benzaldehyde ensembles show some variation. The literature values of the free energies of formation for liquid

Figure 1 water and liquid benzaldehyde at 298 K are as follows: -56.687 and +2.24 kcal/mol. It appears that AM-1 does not predict the free energy of formation for the condensed phases well on an absolute basis.

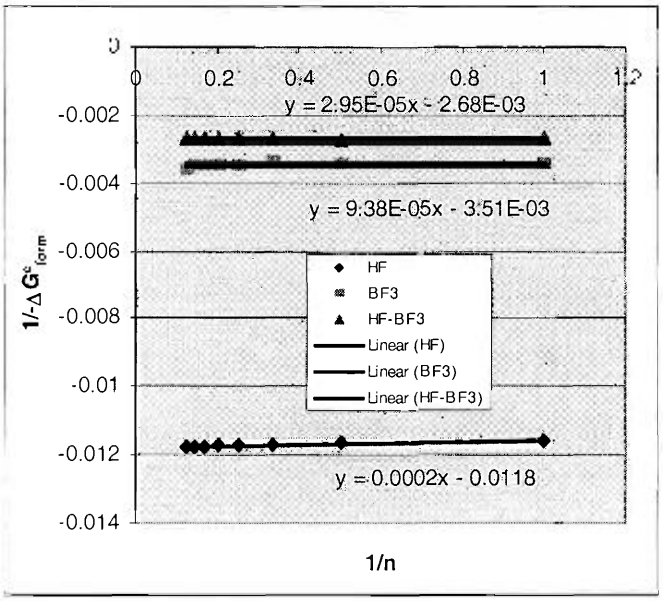


HF/BF_3 System. The HF/BF_3 system has been used to formylate toluene in a commercial process by which the conversion agent, HF/BF_3, is easily recycled. It appears to be prudent to model this system so as to discover how it does permit the breaking of the complex with *p*-tolualdehyde.

We used the approach described above to estimate the thermochemical properties of HF, BF_3, and a mixture of HF/BF_3 at 298 K (Table 2).

Table 2 Free Energies of Formation for Ensembles of HF, BF_3, and HF/BF_3 (1/1 molar), kcal/mol

	1	2	3	4	5	6	7	8	liquid
HF	-86.5088	-85.8145	-85.7029	-85.398	-85.6444	-85.1554	-85.1165	-85.034	-84.7458
BF_3	-290.394	-289.725	-294.413	-287.777	-288.179	-286.461	-286.843	-277.216	-284.9
HF-BF_3	-379.027	-371.954	-377.644	-373.602	-373.382	-372.06	-377.247	-375.193	-373.134



We may compare the predicted values of the free energies for the isolated molecules with literature values of the gas at 298 K. For HF, the predicted and observed values are –86.5 and –66.35 kcal/mol; whereas the predicted and observed values for BF_3 are –290.34 and –54.84 kcal/mol. A plot of these data for $1/-\Delta G^\circ_{form}$ *versus* 1/n is shown in Figure 3. Notice how the calculated free energies of formation do permit an estimate of the free energy of formation for the species as a liquid (*i. e.*, as $n \rightarrow \infty$; $\Delta G^\circ_{form,\, n} \rightarrow \Delta G^\circ_{form,\, liquid}$. The extrapolated free energies of formation for HF and BF_3 Figure 3 condensed phases are –84.7 and –284.9 kcal/mol. Since HF is not a liquid at 298 K, this predicted value of the free energy represents the energy of a hypothetical state and thus no comparison can be made with literature values for liquid HF. The literature value[1] for the free energy of liquid BF_3 at 298 K is –57.5 kcal/mol. The AM-1 program clearly missed predicting accurately the free energy of liquid BF_3 at 298 K.

Consider now the free energy of solvation at 298 K for one mole of HF placed into one mole of BF_3 to form one mole of mixture using the extrapolated values for the free energies of formation: -373.134

– (-84.7458 – 284.9) = -3.49 kcal/mol. This negative free for the solvation suggests that the two species will mix spontaneously. The heat of mixing, not shown here, for this solvation is –4.8 kcal/mol.

E. Summary

The simple semi-empirical program AM-1 appears to show mixed results in predicting the enthalpy and free energy of condensed phases. Accurate predictions for the free energies of the condensed phases will be helpful for the successful completion of this project. More work will be needed to correctly model the condensed phases of the model systems.

Appendix III—Proposal to BP-UK Fuels

Task I—Complex Breaking. We propose the following experiments to break the acid/aldehyde complex. A synthetic mixture of *p*-tolualdehyde and the acid will be prepared in the ratio of aldehyde/acid normally encountered at the end of a reaction experiment. For example, with triflic acid, the acid/aldehyde ratio can vary from 2 mol/mol to nearly 10 mol/mol. With the perfluoroalkanesulfonic acids described in the Exxon patents, the acid/aldehyde ratio is closer to 1 mol/mol. Vacuum is usually drawn above the liquid so as to facilitate complex breaking and we will do the same here. Exxon inventors describe the use of wiped surface evaporators to assist and we plan to use a Rotovap to simulate the effect of the wiped surface evaporator. With these ideas in mind we propose the following experiments.

Experiments 1-3. A synthetic mixture of triflic acid and *p*-tolualdehyde (2 mol/mol) will be heated in a Rotovap at 100°C under a vacuum of 250 milli-Torr with the condenser operated at –20°C using refrigerated ethylene glycol as the cooling media. The vapors will pass through a product receiver held at dry ice-acetone temperature. In separate experiments this synthetic product mixture will be treated with: a) dichloromethane and diethyl ether (equimolar mixture, total moles/mole of acid = 2/1); b) diphenyl ether (equimolar to the acid); and c) trifluoroacetic acid (equimolar to the acid).

Experiments 4-6. The conditions of experiments 1-3 will be repeated for breaking the complex formed from a synthetic mixture of perfluorobutanesulfonic acid and *p*-tolualdehyde (1 mol/mol).

From these six experiments we expect to learn if the modeling results described in Appendix II can be confirmed and to determine if the procedure for recovering benzaldehyde from chloroaluminate IL's that was reported in a recently reviewed manuscript can be applied to these perfluoroalkanesulfonic acid systems.

If the experiments 1-3 are partially or totally successful, we suggest additional experiments for which the triflic acid/aldehyde ratio is 10 mol/mol using the most successful of the complex-breaking agents revealed in the first three experiments.

Task II—Reaction with and Reuse of Perfluoroalkanesulfonic Acids. We will attempt to reproduce examples 1-3, and 5 of US Patent Application 20002068841 (Perfluoroalkanesulfonic acid—Saleh) as the primary means to complete this task. One purpose of this task is to determine if the examples in the patent can be reproduced in our laboratories. Secondly, we will endeavor to improve upon the technology revealed in these patent examples using our experience acquired with triflic acid and ionic liquids. The following is an abbreviated summary of these four examples showing reaction conditions, yields and separation protocol.

Example 1: Perfluorooctanesulfonic acid (8 g, 17 mmol) and toluene (3 mL, 26 mmol) were placed in an autoclave which was pressurized with CO to 1050 psig, heated to 50°C for 2 hours with stirring. The reaction was quenched by pouring over ice water,

and the organic layer was extracted with diethyl ether. GC analysis showed about 2% conversion of the toluene with 93% regio-selectivity to *p*-tolualdehyde and 7% to *o*-tolualdehyde.

Example 2: The procedures of Example 1 were used except that 6 g of perfluorohexane sulfonic acid (16 mmol) was added to the reactor along with the toluene. The toluene conversion was 2% and the regio-selectivity was the same as example 1.

Example 3: The procedure of example 1 was repeated except that 7 g of perfluoroethoxyethanesulfonic acid (22 mmol) was added to the reactor along with the toluene. The toluene conversion was 9% and the regio-selectivity was the same as example 1.

Example 5: The complex of perfluorohexanesulfonic acid with *p*-tolualdehyde (26 g. 0.05 mol each of acid and aldehyde) was dissolved in trifluoroacetic acid (53 g, 0.46 mol). This solution was fed to a wiped film evaporator over a period of 20 min. The evaporator was run with a wall temperature of 100°C under a pressure of 0.25 mm Hg. The vapor was collected on a cold finger. This liquid was *p*-tolualdehyde. Perfluorohexanesulfonic acid was not volatilized whereas the trifluoroacetic acid was volatilized and collected in a dry ice trap.

These patent examples are bereft of details which we must provide before meaningful experiments can be completed. We propose to use as control experiments reactions involving toluene and triflic acid. The patent examples described above use acid/substrate ratios of ~ 0.67 mol acid/1 mol of toluene to 0.84 mol acid/mol of toluene. In our experience with triflic acid, such low acid/substrate ratios result in low conversions of toluene and for reaction periods longer than ½ hour, we notice the formation of the condensation products: di- and tri-tolymethane. Thus, we propose to add experiments to the list of the three patent examples shown above which explores the effect of increasing the perfluoroalkanesulfonic acid/substrate ratio to ~ 2 mol/mol. The purpose of these additional experiments is to determine if higher yields can be obtained.

Proposed experiments. The following experiments are proposed to confirm the Exxon patent examples and to extend our knowledge base for this technology.

Experiments 1-9. Conduct high-pressure carbonylation of toluene at 1000 psig and 50°C in a Parr autoclave. In separate experiments combine acid (triflic acid control; perfluorobutanesulfonic; or perfluoroethoxyethanesulfonic acid) and substrate in the ratio of 0.67 mol acid/mol substrate for reaction times of ½ h, 1 h, and 2 h.

Experiments 10-18. Repeat these experiments using an acid/substrate ratio of 5 mol/mol.

The information gained from these experiments will allow us to know the effect of reaction time and and acid/substrate ratio upon the reactivity and the regioselectivity of the systems.

Task III Toluene carbonylation over zeolites. We will attempt to synthesize a copper-exchanged ZSM-5 catalyst after the examples described in US 3,369,048. ZSM-5 is chosen over X or Y zeolites described in this patent because it is known to demonstrate a higher acidity than either of these zeolites. Whereas, ion exchange is usually conducted using aqueous solutions of metal salts, we reported a method to introduce Cu(II) into ZSM-5 as highly dispersed, divalent ions thus precluding the problems of overexchange often reported for Cu/ZSM-5 zeolites.[108] The zeolite is contacted with a solution of Cu(II)$acac_2$ dissolved into acetonitrile under reflux finally separated by filtration. The solvent is removed by vacuum. The following experiments are suggested:

Experiment 1: Five grams of ZMS-5 (SiO_2/Al_2O_3 = 70; 300 millimoles H^+/g), without Cu, is heated to 325°C in an autoclave and hydrogen gas is passed over the solid to remove all of the water. This solid is cooled to room temperature and toluene (0.2 g dissolved in 10 mL of decane) is introduced into the autoclave. Dry CO and dry HCl are put into the autoclave in equimolar amounts to a final pressure of 500 psig. The pressure is maintained at this level until the pressure no longer changes with time. The gas is vented and the tolualdehyde is released from the solid by treating with steam. After separating the liquid from the solid, it is heated to 325°C in inert gas to regenerate it.

Experiment 2: Five grams of ZSM-5 (SiO_2/Al_2O_3 = 70; 300 millimoles H^+/g) exchanged with Cu(II) at 50% level (150 millimoles Cu/g, 1.5 wt% Cu) is heated to 325°C in an autoclave and hydrogen gas is passed over the solid to remove all of the water and reduce the Cu(II) to a lower oxidation state. This solid is cooled to room temperature and toluene (0.2 g dissolved in 10 mL of decane) is introduced into the autoclave. Dry CO and dry HCl are put into the autoclave in equimolar amounts to a final pressure of 500 psig. The pressure is maintained at this level until the pressure no longer changes with time. The gas is vented and the tolualdehyde is released from the solid by treating with steam. After separating the liquid from the solid, it is heated to 325°C in inert gas to regenerate it.

These two experiments will determine if zeolites can be used to carbonylate toluene, and if Cu is necessary for the reactivity/selectively. One additional experiment may be desirable if either of these two experiments is successful. This experiment will attempt to reuse the catalyst as outlined below.

Experiment 3: The more active catalyst of experiments 1 and 2 will be reacted again by first reducing it with heating in H_2 to 325°C, then cooling to room temperature, and adding 10 mL of a substrate (0.2 g) solvent mixture to the catalyst in the autoclave. Dry CO and dry HCl are put into the autoclave in equimolar amounts to a final pressure of 500 psig. The pressure is maintained at this level until the pressure no longer changes with time. The gas is vented and the tolualdehyde is released from the solid by treating with steam. After separating the liquid from the solid, it is heated to 325°C in inert gas to regenerate it. This protocol is repeated three times to determine the activity/regioselectivity as a function of reuse.

Task IV—High risk, high return technology. Hexanitrato ceric acid is a powerful oxidizing acid that has been extensively studied[109] in the nuclear industry for the dissolution of refractory actinides and the oxidation[110,111] of organics. It is different from the other super acids that have been studied for several reasons:

1. It can be regenerated electrochemically and thereby would be an ideal replacement for potassium persulfate as an oxidant.
2. In nitric acid solutions, cerium serves as a catalyst and it is nitric acid that is reduced to liberate nitric oxides which can then be recovered and recycled by off gas scrubbing and evaporation.
3. Ceric acid requires glass-lined vessels or titanium, but this acid is easily reduced to cerous nitrate using oxalic acid. The cerous ion in nitric acid is compatible with stainless steel.
4. Cerous nitrate is easily recycled using liquid/liquid extraction.
5. The acid works well in aqueous mixtures. With excess air additions, toluene is readily oxidized to carbon dioxide. This would be a reaction that we would avoid by using an atmosphere devoid of oxygen and containing only CO.
6. We find no literature discussing the use of this acid for arene formylation, but we speculate that it may be a suitable substitute for potassium persulfate to initiate radical reactions.
7. If arene carbonylation can be conducted in aqueous media, we suspect that the presence of water will also dissociate the aldehyde and acid complex.
8. If the desired reactions can be catalyzed using hexanitrato ceric acid, then cerium may provide a system where the catalyst is regenerated electrochemically and recycled using stainless steel equipment.
9. We note that nitric acid chemistry, while it has environmental problems, is generally less objectionable than fluoride-based chemistry.
10. We find no applicable patents discussing this chemistry.

For these reasons we recommend that ceric acid chemistry be evaluated for this application.

It is proposed that we conduct one set of reactions to determine if the Ce-based system can affect the carbonylation of toluene in the presence of CO at 150 psig and room temperature. The result of this reaction in terms of activity and selectivity will dictate further work in this area. If the reaction rates are not suitable, then one more set of experiments are contemplated for which the temperature and/or pressure can be adjusted accordingly.

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